INTRODUCTORY
CHEMISTRY

FOR

INTERMEDIATE SCHOOLS

BY

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PREFACE

THE object of a school course of Chemistry is not primarily to impart a number of chemical facts; but, rather, to develop scientific methods of work, to promote thought, and to encourage reasoning. In the scientific work which legitimately finds a place in the curriculum of a secondary school, every experiment is a link in a continuous chain of inquiry. From a simple fact, as for instance the rusting of iron, it is possible to proceed step by step to the experimental examination of more complex changes and so to determine important principles of chemical science. From beginning to end the practical work must have a definite purpose in view.

Such appears to be the spirit in which the suggestive syllabus of the Department of Technical Instruction for Ireland, adopted by the Irish Intermediate Education Board, has been drawn up, and constant reference has been made to it during the preparation of this book.

In the past, the practical work which young students of chemistry have been set to perform has had little, if any, connection with his lessons in the lecture room. Fortunately, it is now beginning to be recognised that only the knowledge of chemical theory which is the immediate outcome of his own experiments is of real value to the beginner. For this reason, what have so long been known as "practical" and "theoretical" chemistry have not been separated, and an attempt has been made to train the pupil to understand that chemistry is a science

in which no statement without an experimental basis is permissible.

Though the particular needs of Irish schools have in a measure decided the contents of the book, teachers will find that the course described is suitable for any secondary school as an introduction to the study of chemistry. Public examining bodies are year by year showing sympathy with the "research" method of teaching advocated by so many men of science who desire to increase the educational value of the teaching of science in schools.

My hearty thanks are due to Prof. R. A. Gregory and Mr. A. T. Simmons for permission, of which I have fully availed myself, to make free use of any of their published books and diagrams, and for invaluable suggestions and assistance during the preparation of the MS. and the passage of the book through the press.

<div align="right">L. M. JONES.</div>

August, 1902.

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CHAPTER I.

MATTER AND PROPERTIES OF MATTER.

BEFORE beginning the study of Chemistry the student must become familiar with the properties of **Matter,** *i.e.* the properties which matter may possess. The word matter is so frequently used that it is necessary to form a definite idea of what it means. We may say that *anything which possesses weight is a form of matter.* Thus, water is matter ; so is wood ; but water is not the same kind of matter as wood. Air also possesses weight and is matter. It may be said that all the various things, or substances, which you can handle are material things.

Heat is not matter, for it has no weight ; a hot body does not weigh any more than the same body when cold.

It is clear, then, that there are very many different kinds of matter, and, before studying further, we must try to classify some of the properties which they may possess, for probably many kinds of matter will resemble each other in some particulars, that is will have some property in common.

1. HARDNESS OF MATTER.

i. **Measurement of hardness.**—Procure pieces of flint, rock-crystal, a tumbler, chalk, lead, pocket-knife, iron, copper, brass, wood, soap, wax, a turnip, carrot, potato, or apple.

(*a*) Notice that these things differ from one another, and consider in what ways they are different. They differ in hardness, shape, size, and colour.

(*b*) Select one of the things, and notice that it will scratch some substances but not others. Test the things which the

knife will scratch, or cut, and the things it will not cut.   Test in
the same way the things the finger-nail will scratch and those
it will not scratch.

(c) Arrange the substances in pairs as below, so that one is
scratched by the other.   In this way a continuous table in
which the substances are arranged according to their hardness
can be drawn up, thus :

Flint	scratches	glass.	Copper	scratches	lead.
Glass	„	iron.	Lead	„	chalk.
Iron	„	copper.	Chalk	„	wax.

Things differ in hardness.—If you were asked to say how
the things provided differ from one another, you would
probably say that they differ in size, shape, colour, hardness,
and in other ways.   Now consider exactly what you mean
by the property of hardness.   A stone is hard, so is a piece
of wood, and so is a piece of iron, but they are not of the
same hardness.   Some things, then, are harder than others.

It is often easy to decide which is the harder of two
things.   For instance, you know that a knife is harder
than a piece of wood ; for you can often dig your thumb-
or finger-nail into the wood, but you cannot dig your
nail into a steel knife.   Also, you can cut wood with a
knife, but you cannot cut it with a piece of india-rubber,
because the india-rubber is softer than the wood.   All
things which a knife will cut, or scratch, are softer than the
knife, and all things which it will not cut, or scratch, are
harder than it.

In the same way, things like potatoes, some woods, chalk,
bread, blotting paper, and soap can be scratched by the
finger-nail, and are therefore softer than the finger-nail.
Things like iron, glass, and flint cannot be scratched, or cut,
by the finger-nail, and are therefore harder than it.

The test of hardness.—You will now understand the
way to find out which is the harder of two things.   What
has to be done is to test which will scratch, or cut, the other.
If you were asked whether glass or flint was the harder, you
should try if the flint will scratch the glass.   It does.   Will

the glass scratch the flint? It will not. The flint is consequently the harder.

In the same way, if you were given a large number of different things and told to arrange them in the order of their hardness, you would take any one of the substances and find which of the others it would scratch and which it would not scratch. Then another would be taken, and the same tests made, and so a list like the one below would be made. This is the method always adopted to find out if one thing is harder than another.

1. Diamond.	5. Iron.
2. Rock-crystal.	6. Copper.
3. Glass.	7. Lead.
4. Steel.	8. Wax.

The hardest substance is first in the list, the next hardest is second, and the softest is last. Any of the substances will scratch a substance lower in the list, and can be scratched by substances higher in the list. Diamond is seen to be the hardest substance; it will scratch every other thing. Emery is also very hard, and is therefore used for polishing many things.

2. SOLIDS, LIQUIDS, AND GASES.

i. **Three forms of matter.**—In addition to some of the solids used in the last exercise, take a tea-cup, tumbler, salad-cream bottle, round medicine phial, piece of india-rubber tubing, and a large basin, or pan, of water. Also obtain water, milk, quicksilver, and any other liquids available.

FIG. 1.—A liquid takes the shape of the containing vessel.

(a) Notice that the solid things upon the table are of different shapes, and that the shapes do not alter.

(b) Show by pouring the same amount of water or other liquid into different vessels that the shape of the water depends upon the shape of the vessel. (Fig. 1.)
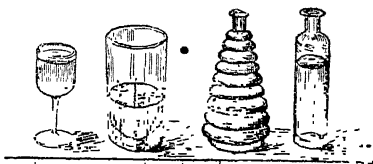
(*c*) Collect a bottle of ordinary gas, and use it to show that a gas has no surface and spreads itself through as much space as it can. The following is a way to do this :

Fill a bottle with water, and invert it in a basin of water ; then displace the liquid with gas led from a jet by a piece of india-rubber tubing (Fig. 2). Now insert a cork into the neck of the bottle while it is still under water, or cover the mouth with a glass plate, and lift the bottle out of the water and place it on the table. The gas has the size and shape of the bottle.

Open the bottle and wave it about ; you immediately notice the smell of gas throughout the room, and know from this that the gas is everywhere in the room, and therefore has the size and shape of the room.
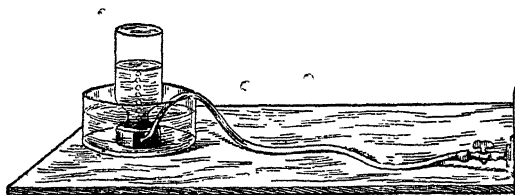

FIG. 2.—Collection of a bottle of ordinary coal gas.

Solids.—Most of the things you see around you have a certain size and shape of their own. The table in front of you and the desk you sit on have the same shape now as when they were first brought into the school, if no one has done anything to them. In the same way, a stone, a brick, a piece of india-rubber, or a tumbler keep their own shape unless someone breaks them. Things of this kind, which have a size and shape of their own, and remain of the same size and shape so long as they are not interfered with, are called **solids**. Some solids are harder than others, and some can have their shape altered more easily than others. But none of them change by themselves. If you place a tumbler upon a table, and leave it for a while, you expect to find it there when you come back, and not changed into a bottle, because you know the shapes of solids do not alter unless someone alters them.

Liquids.—If you put a stone into a tea-cup, then into a tumbler, and then into a basin, you know that the size and shape of the stone remain the same in all the vessels. This is also true of any solid. But if a certain amount of water, say a wine-glassful, is taken and poured into a tumbler, you know that the water will not keep the same shape that it had at first. The water could be poured successively into vessels of different sizes and shapes, and finally into the wine-glass again; but though *the shape would keep on altering*, being in every case the shape of the containing vessel, *the size would remain unaltered.* Things which behave in this way are called **liquids.** Such as, for instance, vinegar, oil, milk, beer, and lemonade.

There are other properties of liquids with which you are familiar. In the first place, the surface of a liquid at rest is always horizontal. You may shake the liquid up into a heap by jerking the vessel containing it, but as soon as you leave it alone it settles down again until its surface is horizontal. The fact that the liquids can be poured from one vessel to another shows another property, namely, that liquids flow. A liquid can also be broken up into small round drops, such as the drops of water which form rain.

Gases.—When a gas is spoken of, you think of the gas used to illuminate rooms and streets. There are, however, many other gases, and you will learn about them later. Some gases possess a distinct odour and some do not ; some are poisonous and some are harmless ; some, like coal-gas, will burn, and some will not. But all gases are alike in this respect ; they spread out and fill completely the vessel which holds them. A bottle of gas cannot be kept unless it is tightly corked, for after a while the gas escapes into the surrounding air.

Comparison of the size and shape of solids, liquids, and gases.—Solids, liquids, and gases are not alike as regards size and shape. Solids have a size and shape of their own,

which may differ for each solid, but remain the same for one particular solid.

Liquids have a size of their own, but always take the shape of the vessel in which they are contained.

Gases have no definite size or shape ; both these properties depend upon the space in which the gases are confined. However small a quantity of gas may be, it always spreads out until something prevents it from taking up more space.

3. THE SAME KIND OF MATTER CAN EXIST IN DIFFERENT STATES.

i. **Ice, water, and steam are different forms of the same kind of matter.**—(*a*) Procure a lump of ice and notice that it is a solid, or has a particular shape of its own, which, as long as the day is sufficiently cold, remains fixed.

(*b*) With a sharp bradawl, or the point of a knife, break the lump of ice into pieces, and put a convenient quantity of them into a beaker. Place the beaker in a warm room, or apply heat from a laboratory burner, or spirit lamp. The ice disappears, and its place is taken by what we call water. Notice the characters of the water are those of a liquid. It has no definite shape, for by tilting the beaker the water can be made to flow about.

(*c*) Replace the beaker over the burner and go on warming it. Soon the water boils, and is converted into vapour, which spreads itself throughout the air in the room, and seems to disappear. The vapour can be made visible by blowing cold air at it, when it becomes white and visible, but is really no longer vapour, for it has condensed into small drops of water.

ii. **Some metals assume the liquid condition when heated.**—Heat a piece of lead, or zinc, in an iron spoon. Observe the solid metal eventually changes into a liquid, and that more heating is required to effect the change than in the case of ice.

iii. **Gradual change of state.**—Warm a lump of sealing wax, or bicycle cement, in an iron spoon, and notice the gradual conversion into a liquid.

Changes of state in the same kind of matter.—To the fact that there are three kinds of material things, we must add another idea, viz., that the *same matter can exist in all three forms*. The change in the state of matter, whether from the

solid to the liquid condition, or from the liquid form to the gaseous state, is most easily brought about by heat. Reverse changes, viz., from gas to liquid and from liquid to solid, can be effected by cooling.

The degree of heating required to bring about the above changes varies very greatly with different substances. Iron must be heated very much more than ice before it can become a liquid. Alcohol, again, has to be cooled to a much greater extent than water before the liquid condition gives place to that of a solid.

When a substance has, as a result of heating, successively passed through the solid, liquid, and gaseous states, then, if the conditions are reversed and the gas is continuously cooled, the liquid form is first reassumed, and subsequent cooling causes the liquid to change again into the original solid.

Sudden and gradual changes.—The circumstances attending the change from the solid to the liquid, or from the liquid to the gaseous state, are not always the same as in the case of water. When a crystal of iodine is heated, it appears to pass suddenly from the condition of a solid to that of gas. Camphor is another instance of this sudden transition from solid to vapour. When, on the other hand, sealing-wax is heated, it very gradually passes into the liquid condition, and may be obtained in a kind of transition stage—neither true solid nor true liquid.

4. NO LOSS OF MATTER DURING CHANGE OF STATE.

i. **Weight in solid and liquid states.**—(*a*) Place a piece of ice in a flask and weigh the two. Notice the weight, then melt the ice by warming the flask, and observe that the weight of the water is practically the same as that of the ice.

ii. Weigh a test-tube containing a little butter or wax. Notice the weight. Melt the butter or wax, and observe there is no change of weight.

Change of state does not imply change of weight.—At first sight it would seem that when the state of a substance is changed its weight also undergoes alteration. This, however, is not the case. If 1 lb. of ice is melted, 1 lb. of water is produced, and if the lb. of water is converted into steam, there is still only 1 lb. of vapour, provided that care is taken not to let any escape. This constancy in the weight of the water, whether in a solid or liquid state, applies equally to all substances.

5. OTHER COMMON PROPERTIES OF MATTER.

i. **Some common things.**—Obtain the following :—Piece of glass, transparent liquids, sealing-wax, roll-sulphur, slate, lead, feathers, cork, sponge, sheet-lead, india-rubber, cane, air-ball or pop-gun. Examine each of the things thoroughly, and consider what you would want to know about it in order to describe it to a person who had never seen it.

(*a*) Notice, for instance, that glass can be seen through (transparent), and breaks easily (brittle). Lead, and most other things, cannot be seen through (opaque). Lift the lead. It is heavier than pieces of the other substances of the same size (dense). Beat it into a thin sheet (malleable). Bend or twist it and it remains in the shape it is made (pliable).

(*b*) Stretch, bend, compress, or twist the india-rubber, and then release it. It goes back to the original shape (elastic). Compress an air-ball or a pneumatic tyre. Release it. It also springs back to the original shape ; hence air is elastic. Bend cane or whalebone (flexible).

Glass.—You have now taken a few common things, and examined them to learn some of the properties things may possess. Consider the properties of **Glass.** You can see through it, and all things you can see through are called **transparent.** Glass, then, is transparent. Some other things which are transparent are rock-crystal, air, water, and many other liquids.

If the plate of glass is dropped, or if it is struck by a hammer, it breaks into pieces. Things which break into fragments in this way are called **brittle.** Some other brittle things are cast-iron, sticks of sealing-wax, roll-sulphur, slate.

In being brittle all these things are like glass. In what way do they all differ from glass? You cannot see through them, or, as is usually said, they are **opaque.**

Lead.—What have you learnt about this substance? You say it is very heavy. But it is not enough to say that, for if you take a sufficient quantity of feathers they will weigh as much as the piece of lead. There is, of course, no difference between the weight of a pound of lead and a pound of feathers. What you meant to say was that a small piece of lead was very heavy. Because of this we speak of lead as being very **compact** or **dense.** Some other dense things are copper, gold, etc. If a very large thing has but a small weight, it is not dense. Such things as cork, sponge, camphor, are not dense.

Lead can be hammered out into sheets. Sheet lead is used for lining tea chests and other boxes. Solids which can be beaten out into sheets are called **malleable.** This property of malleability is also possessed by copper, gold, platinum, but gold is more malleable than any other solid.

Lead can be bent and does not spring back; its shape after bending, or twisting, remains just as you left it. All things which remain just as they are left after bending, or twisting, are called **pliable.** Copper, paper, and sheet-tin are also pliable.

Some substances, however, may be readily bent, but afterwards if left to themselves they again spring back to their original form. Such substances are said to be **flexible,** *e.g.* cane, whalebone, etc.

India-rubber.—The name of the chief property of india-rubber is actually given to it when it is made into long threads. These are called **elastic.** India-rubber is called elastic, because though you pull it, squeeze it, or bend it, it returns to its original shape and size when you leave it alone. It is the property of going back to its first shape and size after being forced out of it that is called **elasticity.**

If you squeeze an air-ball, or push upon a pneumatic tyre, when you remove your fingers they spring back to the shape they had at first. These experiments tell you that air is very elastic.

6. OTHER COMMON THINGS.

Obtain :—a sponge, clean white blotting-paper, glass-funnel, a cane, crystals of rock salt and sugar candy, powdered salt and sugar, several crystalline substances, such as washing-soda, borax, and rock-crystal, some flour and soot, matches, magnesium ribbon, a taper, and some clay.

i. Notice the holes or pores in sponge (porous). Place a sponge in a saucer of water, and notice that the water disappears ; it goes into the pores in the sponge, and can be squeezed out again. Filter water through blotting-paper, folded to form a cone, and placed in a funnel (Fig. 3). Show that clay will hold water (impervious).

ii. Observe the regular shape of particles of salt and sugar (crystalline). Particles of flour and soot have no regular shape (amorphous).

iii. (a) Add salt or sugar to water, and notice that it disappears (dissolves, or is soluble).
(b) Show that sand, slate-pencil, and many other substances are insoluble.

iv. Burn a match, paper, magnesium ribbon, etc. (combustible). Find some things which will not burn (incombustible).

Sponge.—By squeezing a sponge you at once find out that it is elastic. But the first fact you notice about the sponge is that it has lots of holes in it. The holes in a cane are smaller than those in the sponge. A thing which is full of holes, or pores, is called **porous**. The sponge, therefore, is very porous. But for a thing to be porous it is not necessary to be able to see the holes.

If you take a piece of blotting-paper, and fold it as shown in Fig. 3, put it into a glass funnel, and pour some water on it, the water finds its way through the paper because, though you cannot see them, there are holes in the paper, or it, too, is porous.

Even iron is porous to a small extent, and water is some-

times filtered through a certain kind of iron. Pumice, charcoal, sandstone are also porous. Things which will not allow water to pass, or filter, through them are called **impervious**. Clay is an impervious thing; so is glass and india-rubber.
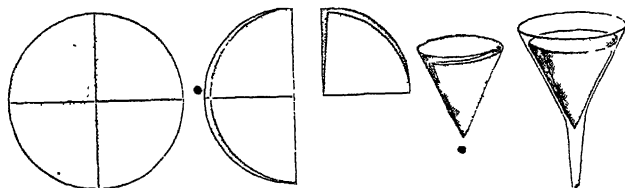

FIG. 3.—How to fold a filter paper

Salt and sugar.—Salt and sugar are **crystalline** substances, that is, each small piece has a certain regular shape, which for a particular substance is always the same. Solids which take a regular shape like this are called **crystals**. Other common things which occur in crystals are washing-soda, borax, diamonds and some other precious stones, rock-crystal, and many other substances. It is not always possible to recognise the shape of the crystals, as they are seldom perfect; we usually recognise a crystalline substance by the bright, glassy appearance of the particle.

Things which are not crystalline are called **amorphous**, a word which is made up of two Greek words meaning 'without shape.' Flour, soot, etc., are examples of amorphous things.

Now, if you put the salt, or sugar, into a glass, or bottle, of water, and shake, or stir, the water, you will notice that the salt, or sugar, disappears, and if you taste the water you easily recognise the presence of the salt, or sugar. You may say the salt, or sugar, has **dissolved**, or is **soluble**. Washing-soda, nitre, and borax dissolve in water.

Substances which will not dissolve in water are said to be

insoluble. Many things are insoluble in water; for instance, sand, gravel, slate-pencil, coal, chalk. However long you leave these in water they will not dissolve.

Things which burn.—Many things easily burn in the air when made hot enough. If you hold a match in a gas-flame you can easily show that it continues to burn after taking it out. Similarly, you can make a piece of magnesium ribbon burn. Tapers and pieces of paper also burn quite easily. Things which burn in this way are said to be **combustible.** Coal, coke, coal-gas, tallow, wood are common combustible things Those substances which will not burn are called **incombustible.** Slate, iron, bricks, glass, etc., will not burn, and are therefore called incombustible things.

QUESTIONS ON CHAP. I.

1. What have you learnt about the meaning of ' matter ' ?
2. State in your own words what a solid is, what a liquid is, and what a gas is.
3. How do solids and liquids differ from one another ?
4. In what important respects do liquids and gases differ ?
5. What experiment would you do to show that while the size of a liquid remains the same its shape can keep on altering ?
6. What plan should you follow if you wished to arrange three substances in the order of their hardness ?
7. Which is the hardest thing you know ? What is it used for ?
8. Many substances allow water to pass through them very readily. Why is this ? What experiment would you perform to show this in the case of blotting-paper ?
9. What things are said to be soluble ? Name six soluble things.
10. What do you mean by a transparent substance and what by an opaque thing ? Name six transparent things and six opaque things.
11. Explain as carefully as you can what you understand by a dense substance. Name several dense substances.
12. What things are said to be malleable ? Write down the names of as many such things as you can.
13. Why are india-rubber, air, whalebone, etc., said to be elastic ? How would you show that air is elastic ?

CHAPTER II.

IT is now desirable that the student should familiarise himself with some of the substances, or forms of matter, which are of common use in chemical laboratories. He should learn to examine and describe any substance which may be given to him, so as to write out an account of its more obvious properties.

In thus describing a solid it must be remembered that nothing should be taken for granted ; it must be assumed that the person for whom the description is written is, to begin with, completely ignorant of the substance in question, and the observations recorded should be such that the reader is able to form an accurate idea of the substance.

Order of examination.—During the examination of a substance therefore the following points should be noted :

(*a*) State all you can learn of the substance by aid of the sense of sight, that is, state whether it is a solid, liquid, or gas ; whether it is in large or in small lumps, or in the form of a powder. Note also whether the separate particles of the substance have a glassy appearance, with straight edges where the surfaces cut one another; or whether they are dull and lustreless, with no such straight edges. In the first case the substance is probably *crystalline*, in the latter *amorphous*.

If the substance is in large pieces, state whether it is *transparent* or *opaque*. State its colour, and if there are any

other points concerning it which are evident to the eye, be sure to state them.

(*b*) What is learnt concerning the substance by the use of the sense of smell? Has the substance in question any odour? If so, is the odour pleasant or otherwise?

It is very difficult to describe odours exactly; but, if the odour seems sharp, making you draw back and cough, it may be termed a *sharp* or *pungent* odour; if it appears to hang at the back of your throat and prevent you breathing properly, it may be termed a *heavy, choking* odour.

(*c*) The *feel* of the substance should be stated, whether it feels *dry* or *moist*, etc. Some substances have a smooth, soapy feel. Others feel soft, like flour; or gritty, like powdered glass.

(*d*) The substance should then be subjected to other tests. It should be hammered on an anvil or a slab of metal, in order to see whether it is malleable, or brittle. You should test whether it is hard or soft, and, if possible, test about what is its degree of hardness according to the table on p. 3. If you can, you should test whether it is flexible or not, and whether or not it appears to be elastic.

(*e*) These observations having been made, you should, **if permitted by your teacher,** but **never without obtaining such permission,** taste the substance to see whether any definite taste is noticed, such, for example, as a *salt* or *saline* taste; a *sweet* or an *acid* taste.

You may then proceed to test the substance in other ways.

(*f*) Determine whether it is soluble or not in water. In doing this you should also notice whether the addition of water to larger quantities of the substance has any effect. For example, whether the water appears to be absorbed, to generate heat, or to alter the form of the substance, etc.

(*g*) The substance should then be *heated* in a test tube, or in a crucible, and the effects noted. Particular attention

must be paid to changes apparent in the substance itself, such as change of form, change of colour, melting, etc. Any other changes, such as the evolution of vapours, etc., must also be recorded.

For the purpose of heating substances in a chemical laboratory a form of gas burner, called a Bunsen burner, is generally employed (see Fig. 75). The noticeable point about the burner is that near its base are two small holes, which may be covered up by turning a small outer tube. When these holes are closed the gas burns with a bright luminous flame which, however, deposits soot on objects placed in it. When the holes are open the flame is not luminous, and no soot is deposited, and *this non-luminous flame is always employed for heating purposes*.

Sometimes the gas catches alight at the bottom of the burner; it then burns with a long, straight, narrow flame, and not only makes the whole burner inconveniently hot, but also gives out an unpleasant odour. The burner is then said to have *struck back*. When this is the case the gas should be turned off and the burner re-lighted; it may also be remedied by sharply striking the india-rubber tubing with the edge of the hand.

7. EXAMINATION OF SALT.

i. **Salt.**—Examine, according to the scheme just explained, the specimen of common salt provided, and record carefully all your observations. Examine the appearance of the particles to the eye, then employ a lens, and describe what you see. Has it any odour?

ii. Take a little of the substance in your fingers and rub it between them, noting whether it feels gritty or not. Test it to see whether the separate pieces are malleable or not. Taste it and note the result.

iii. Place a *few grains* in a test tube and add distilled water Shake vigorously and notice whether or not the grains disappear. Taste the resulting liquid. Heat carefully for some time in a crucible a small amount of the substance.

A description of common salt.—You will have observed from these tests that salt is a white, crystalline powder. Its separate grains consist of small rectangular solids. They

belong, in fact, to what is known as the cubical system, and natural crystals of considerable size may be obtained (Fig. 4). The natural form of salt, known as *rock salt*, is largely obtained from certain salt deposits which are found, in layers of varying thickness, in Austria, and in England in Cheshire. From these deposits the salt is obtained either by quarrying, or by forcing in water,

FIG. 4.—Crystal of common salt.

pumping out the solution, and then evaporating. It is also
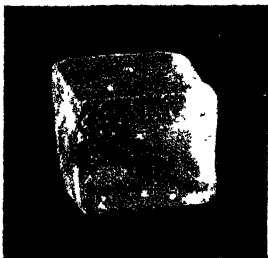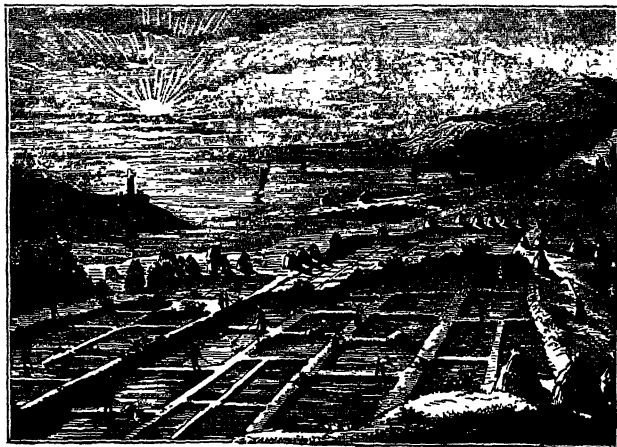

FIG. 5.—Evaporation of sea water.

obtained by the evaporation of sea water (Fig. 5). The natural crystals, which are larger than those of the purified

product employed for household use, are also usually coloured owing to the presence in them of impurities, chiefly iron compounds. Salt feels gritty to the touch, as, indeed, do most crystalline powders ; it is also brittle, the larger crystals readily breaking up into smaller parts when struck.

Salt is inodorous, but has a very sharp saline taste. It dissolves readily in water, the resulting solution possessing a taste similar to that of the solid.

When salt is heated a crackling sound is heard, but no vapour of any description is given off. This crackling, termed *decrepitation*, is due to the breaking up of the larger crystals, under the influence of heat, into smaller portions. If more powerfully heated it may be noticed that salt fuses, but there appears to be no real change in the nature of the substance, since it still possesses the same taste and properties.

It is understood that all the substances dealt with in the following pages are to be examined in a similar manner. In certain cases additional experiments may be required, and these will be noted in their proper places.

8. EXAMINATION OF SAL-AMMONIAC.

i. Examine a specimen of sal-ammoniac according to the plan followed in the case of common salt. Carefully describe all the observations made.

ii. **Action of heat upon sal-ammoniac.**—Place a small quantity of the substance in a crucible and heat. Note the dense white fumes which rise. Place a large funnel over the heated crucible and observe the result. Scrape off the funnel some of the white deposit and taste it. Try the effect of water on a part of the deposit, also try the effect of heat upon another portion.

Description of sal-ammoniac.—Like salt, sal-ammoniac is a white crystalline powder. The form of the crystals cannot usually be made out, even with a lens. But, speaking generally, the crystals appear to be feather-like forms.

J.C.   B

If, however, any good crystals are present in the specimens
given to you, the crystals may be seen to resemble the
diagram given (Fig. 6). Sal-ammoniac appears gritty to the
touch, and is brittle. It possesses no odour, but has a very

sharp taste resembling that of salt,
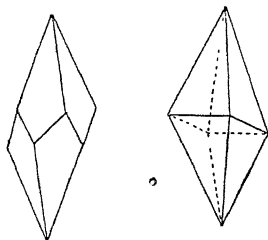but being more acrid, or burning.
It is readily soluble in water.

When heated sal-ammoniac.
soon gives off white fumes, and
eventually completely disappears,
being converted into a gas with-
out first melting to a liquid. In
the cooler air, and upon the
funnel, the gas again forms the

FIG. 6.—Crystals of sal-ammoniac.

solid, for your experiments have shown that the white deposit
is still sal-ammoniac. This phenomenon of the change of a
solid into a vapour, and its subsequent cooling to a solid of
the original nature, is termed **sublimation**, and a few other
examples will probably be met with during your chemical
studies.

9. EXAMINATION OF CHALK.

i. Examine the chalk provided according to the scheme given
on p. 13.

ii. **Chalk is not soluble in water.**—Shake up a small quantity
of chalk with distilled water and allow the liquid to settle.
Pour off a quantity of the *clear liquid* into an evaporating basin
and heat the basin on a piece of wire-gauze over a Bunsen burner,
so that the water is driven off. Notice that no white powder
remains. The chalk was not dissolved, or chalk is insoluble in
water.

Description of chalk.—Chalk will be found to take the
form of an amorphous powder, or amorphous lumps, but if a
thin slice is examined under a lens of high magnifying
power, or under a microscope, the chalk may be seen to be
composed, to a large extent, of minute shell-like structures

(Fig. 7). These are, in fact, the remains of a minute form of animalcule, which even now lives in sea-water. The fact that the chalk is thus composed of the remains of marine organisms proves that it must have been originally deposited at the sea-bottom. Hence, the chalk hills, which occur in many parts of England, must at one time have formed a portion of an ocean bed.
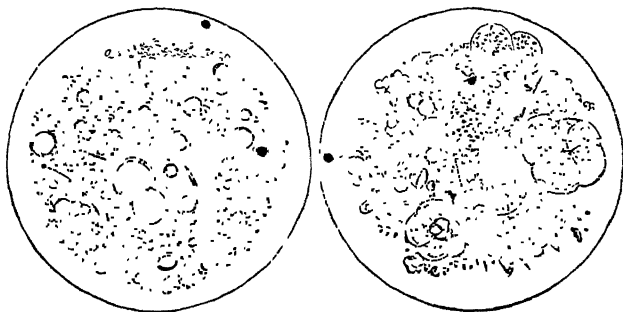

FIG. 7.—Appearance of chalk under a microscope.

Chalk feels softer to the touch than common salt, and leaves a white mark on the fingers. It does not dissolve in water. A few grains placed in a test-tube of water remain, even after vigorous shaking. Heat has *apparently* no effect upon it, the product left in the crucible being still unchanged chalk, while no vapours are observed to be driven off. The student must, however, remember that this is only true when chalk is not subjected to an intense heat, for, as is well-known, it is by strongly heating chalk that the ordinary *lime* of the builder is obtained.

It must be remembered that the substance called 'chalk' used for blackboard writing is usually a composition which is *not* chalk, and must not be used where chalk is described.

10. EXAMINATION OF SAND.

i. **Sand.**—Examine according to the scheme previously employed.

ii. If possible, examine sand under the low power of a microscope.

Description of sand.—Sand is usually obtained as a pale, yellow, granular, crystalline powder. It is very hard, and if a few grains be rubbed over glass, it will be seen that the glass becomes scratched, so that sand grains are harder than glass, and cannot be reduced to a powder by a pestle and mortar. It is quite insoluble in water, and is unchanged by heating. Under the microscope it is observed that each grain has a glassy appearance, but that the edges have been largely worn away by the rubbing of the grains upon one another.

11. EXAMINATION OF BLUE VITRIOL.

i. **Blue vitriol.**—Examine as before.

ii. **The effect of heating blue vitriol.**—(*a*) Place a quantity of powdered blue vitriol in a large test-tube, fitted with a cork through which passes the short arm of a long tube bent at about 60°. The long arm of the bent tube dips down to the bottom of a dry test-tube. Heat the blue vitriol and notice that the vapour driven off collects and condenses in the second tube and forms there a clear, transparent liquid (Fig. 8). When sufficient of the liquid is collected it should be examined. This can be more readily done, if the liquid obtained by several students be collected together for examination.

(*b*) Place a little of the liquid in a test-tube containing also a thermometer. Heat over a very small flame until the liquid boils, and carefully note the boiling point.

(*c*) By means of a graduated pipette take an exact number, say 5 c.c. of the liquid, and deliver it into a clean beaker previously weighed. Weigh the beaker again and so obtain the weight of the 5 c.c. of liquid. From this calculate the weight of 1 c.c., that is the *density* of the liquid.

(*d*) If they are available, mix up some powdered ice and salt in a small beaker and immerse in the mixture a narrow test-tube containing a thermometer and a few c.c. of the liquid. Keep the test-tube in the beaker until the liquid is observed to be

frozen, when the tube should be taken out and the frozen liquid allowed slowly to melt. When melting occurs observe the temperature recorded by the thermometer.
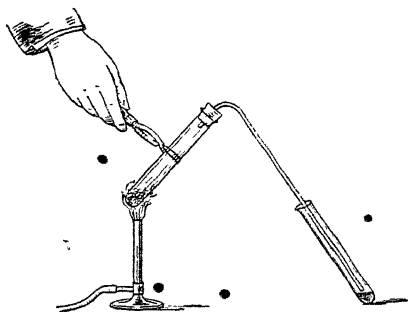

FIG. 8.—Heating blue vitriol

iii. **Change of colour of blue vitriol.**—Drop a little of the white, anhydrous powder (obtained by heating blue vitriol) into water. Notice the instantaneous change to a blue powder with the subsequent dissolution.[1]

Description of blue vitriol.—The blue vitriol with which the student is provided will probably be in the form of tolerably large, blue crystals with plane, glassy faces. If any perfect crystals are present, their shape should be drawn. The crystals will be found to be of the form shown in Fig. 9. They are very brittle, being very readily broken up into a powder when struck by a hammer. The crystals are not so hard as steel, since a knife will easily scratch them, but they

FIG. 9.—Crystal of blue vitriol.

are considerably harder than the finger nails. The crystals are inodorous, but they must *not* be tasted. The crystals readily dissolve in water, forming a blue solution.
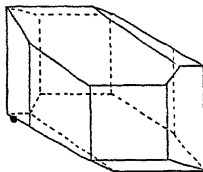
[1] The term *dissolution* is now largely used for the *act of dissolving*, the term **solution** being restricted to the **resulting product.**

When heated in a test-tube the crystals break up into a light blue powder. At the same time a vapour is given off which condenses in drops towards the top where the test-tube is cold. By continued heating the powder becomes almost perfectly white. On allowing to stand this white powder again becomes blue.

If the liquid formed by the condensation of the vapour given off when blue vitriol is heated be collected and examined in the manner described in Expt. 11, ii., it will be found that the liquid is clear and transparent and possesses a boiling point of 100° C. Its density is 1, and its freezing point (or the melting point of its solid form) is 0° C. These tests are sufficient to identify the liquid as *water*. You have consequently learnt that the blue vitriol crystals lose water on heating and form a white powder Many crystals will be found similarly to contain water, and this water is spoken of as **water of crystallisation**. The white powder which no longer contains any water is termed **anhydrous**. A little consideration will suggest that the change back from the white powder to the blue form may be due to an absorption of moisture from the air.

12. EXAMINATION OF GREEN VITRIOL.

i. **Green vitriol.**—Examine as before.

ii. **Effect of heat.**—(*a*) To observe more carefully the effect of heat, place some crystals of green vitriol in a hard glass tube to which is fitted a cork with a long delivery tube bent at right angles. Let the end of this delivery tube dip down to the bottom of a dry test-tube as in Expt. 11, ii. Heat the vitriol and observe that it turns almost white in colour and that a liquid condenses in the test-tube. Remove the test-tube and examine the liquid as in Expt. 11, ii., identifying it as water. Place a little of the light powder in water, observe the green solution which results. Continue the heating of the green vitriol and observe that after some time thick, white fumes are evolved. Now let the delivery tube dip into a second dry test-tube, and continue to heat the green vitriol very strongly. Note the collection of a liquid in the test-tube. Examine the liquid. Note that when heated it forms

thick, white fumes, with a very pungent odour, and that when mixed with a large quantity of water the solution possesses a very sour taste, and that if a drop be placed on some wood it causes it to become charred.

(*b*) Allow the tube of green vitriol to cool and then shake out its contents. Observe that the green vitriol has been converted to a brownish-red powder which does not dissolve in water.

Description of green vitriol.—The green vitriol provided will most probably be found to consist of moderately sized, green crystals, brittle, and soluble in water, forming a green solution. Like blue vitriol, the crystals contain **water of crystallisation** which is driven off by heat, leaving a white powder. This powder, however, is also decomposed by heat, and when strongly heated evolves thick, white fumes which possess an irritating, pungent odour. If these fumes are condensed in a cold test-tube, a liquid is obtained which is obviously not water. It is thicker and more oily than water, and when heated again forms the white fumes; a drop of this liquid in a large quantity of water gives to the water a very sour taste, and the liquid itself is capable of charring wood. It is the substance known as **oil of vitriol.**

The tube is now found to contain a powder differing very considerably from the original green vitriol. It is a brownish-red, insoluble powder, which does not again form the green solution, and is not merely the *anhydrous* green vitriol. Its appearance will be observed to resemble closely that of rust, as, for example, in the stain left if rubbed on paper, etc.

13. EXAMINATION OF SULPHUR.

i. **Sulphur.**—Perform the examination in the manner previously described (p. 13).

ii. **Melting point of sulphur.**—Draw out, in the flame of a laboratory burner, a piece of glass tubing so as to make a small thin-walled tube, about two or three inches long and $\frac{1}{10}$ inch in diameter. Into this tube place some finely powdered sulphur. Tie the filled tube on to a thermometer near its bulb with a

piece of fine platinum wire, and put the thermometer into a beaker of sulphuric acid which has been placed over a burner. (Be very careful not to upset the acid.) Gradually heat the acid and keep it at a uniform temperature by moving the curved stirring rod (shown in Fig. 10) up and down. Notice when the sulphur melts, and at that instant read the thermometer. This reading will be the melting point of the sulphur.
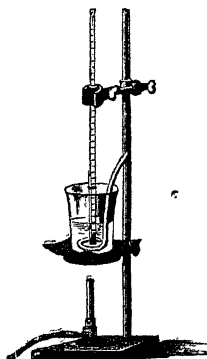
iii. **Effects of heat upon sulphur.**—(*a*) Put some finely powdered sulphur into a large test-tube, using sufficient to fill the tube to a height of about $1\frac{1}{2}$ inches, and heat carefully with a small laboratory burner flame. Occasionally take the tube out of the flame and shake it. When the sulphur has all melted notice that an amber-coloured liquid has been formed. Pour a little of the liquid into a beaker of water. Observe that a lump of yellow sulphur is again formed, which when broken reveals a crystalline structure.

FIG. 10 — Determining of melting point of sulphur

(*b*) Continue to heat the remainder of the liquid sulphur obtained in (*a*) until the liquid boils. Carefully observe the changes in colour and consistency of the liquid. Pour a little of the boiling liquid into cold water. Examine the cooled sulphur, it is *plastic* and not unlike india-rubber.

(*c*) Notice that a yellow deposit has been formed on the cold, upper part of the test-tube in which the sulphur was heated. This is the result of the condensation of sulphur vapour. The deposit is known as *flowers of sulphur*.

(*d*) Place some powdered roll sulphur in a clean, dry, evaporating basin and heat gently on a piece of iron wire gauze. When it has all melted, remove the flame and allow it to cool. As soon as a solid film has formed on the top of the liquid, pierce two holes in it and quickly pour out the remaining liquid sulphur through one of the holes. Remove the film of solid sulphur and examine the yellow, needle-shaped crystals on the sides of the basin (Fig. 11). This kind of crystalline sulphur is known as the *prismatic variety*.

Examine the crystals after a few days. Observe they are now opaque. The prismatic sulphur has changed back again to ordinary sulphur.

(*e*) Heat a little sulphur in a crucible lid or in an iron spoon until it burns. Notice the pale blue flame and observe the pungent smell which the products of the burning possess.

Description of sulphur.—Sulphur forms large, solid masses with a crystalline structure. (It may, however, have been in the form of an amorphous powder.) It is insoluble in water and possesses no smell. When heated it melts and finally
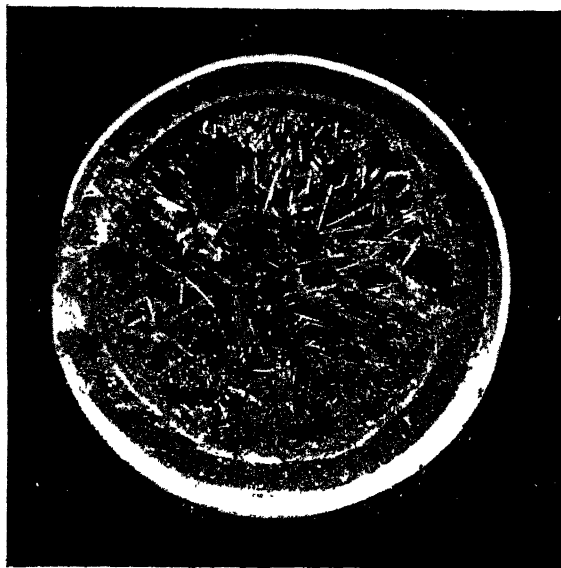

FIG. 11.—Crystals of sulphur   (From a photograph by Mr. H. E. Hadley.)

burns with a blue flame, so pale that it may not be noticed at first, especially in a bright light. Sulphur undergoes a series of changes as it is heated. To follow the changes satisfactorily the heating must be very gradual. When powdered roll-sulphur is heated in a large test-tube it first melts, at about 114° C., into an amber-coloured liquid, which when poured into cold water solidifies into ordinary, yellow sulphur. On continuing to heat the melted sulphur above 114° C., however, it gradually gets darker and darker

in colour, becoming thicker and thicker in consistency, until at about 250° C. it is so viscid that the tube containing it can be inverted and the liquid will not flow. But if the temperature be still further raised, the thick liquid becomes mobile again, and at 440° C. it boils, changing into the vapour of sulphur. The vapour, by sudden cooling, can be changed into a yellow powder, known as **flowers of sulphur.** If the boiling sulphur be poured into cold water it is converted into **plastic sulphur.**

14. EXAMINATION OF NITRE.

i. **Nitre.**—Examine as in previous cases, and record your observations.

Description of nitre —The nitre given you to examine will be most probably in the form of clear, glassy crystals, usually of the form shown in Fig. 12. It may, however, be given to you in the form of a fine, white, crystalline powder. Like other crystals examined it is brittle and gritty to the touch. It readily dissolves in water, and both the solid and its solution in water possess a bitter, cooling, saline taste. When heated, nitre does not give off vapour, and its crystal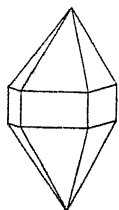s possess no water of crystallisation. Nitre eventually melts if the heating is continued, and on cooling it resolidifies again to a white solid. :

Fig. 12
Crystal of nitre.

Nitre is very largely used for many purposes both domestic and industrial. Thus, it is employed in the manufacture of gunpowder and of fireworks, for medicinal purposes, and for preserving and pickling meat. It occurs very largely in certain hot countries as a deposit on the soil. Bengal, Egypt, Syria, Hungary, and America are the countries from which it is obtained, and although it is now chiefly obtained artificially from other products, yet this natural compound was formerly the chief source of nitre.

15. EXAMINATION OF WASHING SODA.

i. **Washing soda.**—(*a*) Examine as in previous cases.

(*b*) Wash a few crystals of washing soda and dry by filter paper or cloth. Expose to the air for a few days and notice their appearance. You will observe that a white powder forms on the surface on exposure to the air.

FIG. 13.—Fresh crystals of washing-soda. (From a photograph by Mr. H. E Hadley.)

FIG. 14.—Crystals of washing-soda with white deposit. (From a photograph by Mr. H. E. Hadley.)

(*c*) Put some ordinary washing-soda crystals in a test-tube, and in another test-tube put some of the white powder which you have scraped off the surface of some of the crystals, and heat them both. Observe that the soda crystals melt at a much lower temperature than the powder, and although it will be found that water is given off from both, yet there is much less evolved from the powder than from the crystals.

(*d*) Examine the white powder left after heating the crystals, and note that you cannot obtain any more water from it. Observe, however, that the taste of the powder is similar to that of the crystals or of the white efflorescence.

Description of washing soda.—The substance is in the form of clusters of glassy crystals, but in many cases these are covered over with a white powder (Figs. 13, 14). Few perfect crystals are probably present. Their shape, when obtained, is usually as shown in Fig. 15.

Washing soda has a peculiar saline taste, and readily dissolves in water. When heated the crystals very readily melt, and, on further heating, boil, water of crystallisation being driven off.
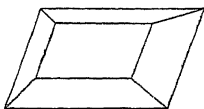

Fig. 15.—Crystal of washing soda.

The formation of the white powder on the surface of the crystals is a subject which deserves further investigation

It appears that the change from the washing-soda crystals to the white powder is due to the loss of some of the water of crystallisation. A similar change frequently happens, but not always, when crystals containing water of crystallisation are exposed to the air, and the crystals are then said to **effloresce.**

Washing soda which contains water of crystallisation readily changes to a white powder containing less water, and this by further heating yields the **anhydrous** compound.

16. SOME COMMON METALS.

i. **Examination of copper.**—(a) Observe and record as before all the properties of copper you can make out.

(b) Heat a strip of copper in a Bunsen flame and record any changes which occur.

(c) Take a piece of copper, which has become grey by being heated, and scratch the surface with a knife. Observe that the bright red lustre is again seen, the greyness being due to a grey deposit, or coating, on the surface of the copper. By renewed heating the copper will again receive a new grey deposit or tarnish.

Description of copper.—Copper is red and has a bright shining appearance, characteristic of many metals, and termed a **metallic lustre.** It has no smell or taste, and is insoluble in water. It is very malleable and pliable. It may in fact be beaten out into very thin leaves, thinner than paper, and it may be drawn out into very fine wire. It is

not as hard as steel, being scratched by a knife. Copper appears to change when heated, for almost directly heat is applied, films of bright colours appear to pass over the surface of the copper, which, after becoming purplish red, finally become a dark grey or black, while the metallic lustre is lost.

[If these changes were not noticed in your experiments, you should again heat the copper and carefully observe them]

You have most probably found the density of copper in your previous work on science. If so, you should look up your physics note-books, and add to your notes on chemistry the value you obtained. It should be about 8·9 gms. per c.c.

NOTE.—Each of the metals now to be examined should be examined according to the scheme given on p. 13, and the observations must be recorded in the note-book.

Description of iron.—Iron generally appears to be a dark grey, or even brown, solid without lustre, but this is because it is generally coated with an outer film, or rust. Clean the iron well with sand paper, and again examine it. You will find that it is a bright, grey metal with a metallic lustre, not as bright, however, as that of copper. It is hard and cannot easily be either hammered out or broken by a hammer. It is insoluble in water.

When heated, iron becomes coated with a dark grey, brittle scale which can be readily detached from the iron. It is also a matter of common experience to all who use their eyes that when iron is heated to redness it becomes malleable and can be hammered into different shapes, whilst it acquires also the power of welding, that is, two pieces of iron when red hot and hammered together **cohere**, or **weld**, into one firm, solid piece. You will doubtless have observed that unused iron, especially when left in a damp place, soon rusts, becoming coated with reddish brown powder or rust,

which, however, may be readily rubbed, or scraped off, again leaving bright metallic iron beneath.

If in a physics course of study you determined the density of iron you should find the value you obtained, and add it to the observations you record in your note-book about iron: it is about 7·8 gms per c.c.

Description of lead.—Lead also appears dull ordinarily; but, when scraped or cut, it is seen to possess a bright lustre, which, however, soon becomes dull. It is malleable and flexible, and has a very high density. Look up the value you have previously found for the density of lead; you will see it is as high as 11·3, lead being one of the most dense of common substances. It is very soft, so soft that you can scratch it with your finger nail and that it can mark paper. It is insoluble in water.

When heated, lead readily melts, forming a liquid with a bright, lustrous surface resembling quick-silver, which soon, however, becomes coated with a scum which re-forms almost as fast as it is scraped off.

The insolubility of lead and the ease with which it can be moulded, bent, and hammered into shape, make it of great use for many industrial purposes, and the student should endeavour to notice how these different properties of the metal are utilised by the plumber, etc., when the metal is used for many of the different purposes to which it is applied.

17. ZINC.

i. **Burning of zinc.**—Take some thin pieces of zinc, and placing them on a piece of fire brick, or on the lid of a crucible, heat strongly with a blow-pipe flame, letting the flame play especially on the edges of the thin sheet. You will observe that the zinc burns in the manner described below.

Description of zinc.—Zinc has a bluish white colour and a metallic lustre. It is, however, usually coated with a very

slight tarnish, though not as marked as that of lead. It is only slightly malleable, and cannot be hammered into very thin sheets without cracking. Its density, which should be determined, if not found already, is about 6·9. It is insoluble in water.

On heating, zinc first melts and then readily gets coated with a scum. If heated strongly, it may get completely converted into this scum, which appears to be of a yellow colour when hot, but becomes dirty white on cooling. If heated sufficiently zinc even burns with a bluish white flame and the production of white fumes.

18. TIN.

i. **The cry of tin.**—If a large piece of tin is available bend it back and fore close to the ear, and notice the peculiar crackling sound which is emitted. This is termed the **cry of tin**; and, although also noticed with other metals which possess a crystalline structure, is not so marked as in the case of tin.

Description of tin.—Tin is a white metal with a very bright lustre. Though not as hard as iron it is considerably harder than lead. If a piece of tin be broken, the crystalline appearance of the fracture is easily noticeable. It is malleable and can be hammered or rolled out into thin foil. The tinfoil used largely in physical laboratories, especially in electrical work, is thus obtained. The density of tin is about 7·3. It is well to note that the so called " tin-saucepans " are not made of tin, but of steel merely *tinned over*, that is covered with a coating of tin. The student will do well to consider wherein such an article possesses advantage over one made solely of iron or steel.

When heated, tin very readily melts—at a lower temperature indeed than lead—and the surface soon becomes coated with a grey film, which when scraped off shows the bright surface of the metal below, which in its turn soon tarnishes, becoming coated with another layer or film. If powerfully

heated in a blowpipe, thin pieces of tin may be made to burn in a manner very similar to zinc.

19. MAGNESIUM.

i. **Density of magnesium.**—Endeavour to determine the density of a few pieces of magnesium ribbon by means of a density bottle, such as you have already used for the determination of the density of small solids. You will probably find, however, great difficulty in getting the magnesium free from air bubbles, which appear to accumulate again after you have got rid of them. This difficulty may be avoided by employing some other liquid such as alcohol ; just as when you had to determine the density of a soluble solid.

Description of Magnesium.—Although not so common a metal as those which have been examined already, magnesium is very largely used in chemical laboratories. It is usually sold in the form of a long, flat, coiled ribbon. It possesses a bright, white lustre, and is malleable and flexible. Magnesium is softer than iron. Its density appears to be low, as the weight of pieces of the metal appears very little

The density is actually about 1·7, being much less than that of the metals you have previously examined.

When heated, magnesium bursts into a brilliant, white flame, giving an intense light, and forming clouds of white smoke which settle on objects held near, while a mass of white ash is left behind.

It is insoluble in water, but the behaviour observed in the determination of the density should be carefully noted.

QUESTIONS ON CHAP. II.

1. Give an account of the points of resemblance and of difference between, sal-ammoniac and salt, copper and lead, blue vitriol and green vitriol.

2. Write a description of chalk, soda, and of sulphur.

3. What is meant by sublimation? Give examples. Explain the meaning of the following terms: efflorescent, decrepitation, deliquescent, lustre, amorphous.

4. How would you ascertain whether a given yellow powder was or was not sulphur?

5. A sphere of lead weighing 18·4 gms. was found to have a volume of 2·3 c.c.   What conclusion would you draw concerning it?

6. How could you readily ascertain (without the application of chemical tests) whether a watch chain consisted of solid gold or was merely . . . . :

7. ·   ·   · would you observe if you heated salt crystals and crystals of soda in different test-tubes?

8. Explain how you would make some crystals of sulphur.

9. How does a crystal of salt differ in shape from a crystal of soda?

10. Why is steam given off when crystals of soda are heated, but not when crystals of sulphur are heated?

11. How would you show that crystals of alum contain water? What is the water called?

12. What do you mean by an efflorescent crystal?   Name one.

13. What happens to a crystal of blue vitriol if it is heated in a test-tube?

14. Describe briefly the effects of heating the following substances in air—green vitriol, lead, sal ammoniac, and wood.

CHAPTER III.

SOLUTION—EVAPORATION—DISTILLATION.

You have already found that one of the ways in which various kinds of matter differ is in their solubility. Thus, though salt, washing soda, sal-ammoniac, and other of the solids you have examined dissolve in water, yet others such as lead, sulphur, and sand do not dissolve, and are said to be **insoluble.** There are many questions which may be asked about solution, and by performing some necessary experiments it is possible to answer them. Such questions, for example, are—Will substances which do not dissolve in water dissolve in other liquids? Do liquids and gases also dissolve in water? If a solid is soluble in water, can any quantity of it be dissolved by a given amount of water? Can the dissolved substance be obtained again from the water? Are substances equally soluble in hot and in cold water? Many other questions could be asked, but those mentioned will serve to show that the simple fact of the solubility in water of some common substances suggests a number of problems which can be answered experimentally and so provide facts to help in future work.

20. SOLUTION OF SOLIDS.

i. **Some solids, insoluble in water, are soluble in other liquids.—** (*a*) *Camphor.*—Stir up camphor with water. Notice that the camphor does not disappear ; it is insoluble in water. Shake up a small lump of camphor with some spirits of wine, in a

small bottle. It gradually disappears, just as sugar does in water.

(*b*) *Sulphur.*—Shake up in a bottle flowers of sulphur with carbon bisulphide, and notice that the sulphur disappears. *Be careful to keep the stopper in the bottle of carbon bisulphide, and do not bring the bottle near a light.*

ii. **Removal of suspended substances by filtration.**—Fold a circular piece of clean white blotting-paper, or a filter-paper, in the manner described in Expt. 6. Insert the folded paper into a glass funnel, and place the funnel into a flask. Make some muddy water by stirring mud into a tumbler of water, or by putting powdered charcoal into it. The mud, or charcoal, remains suspended in the water for a long time.

Pour the muddy water carefully on to the filter-paper in the funnel, and observe that the water which drops through is quite clear. The mud is left on the paper.

iii. **Non-removal of dissolved substances by filtration.**—Similarly, filter a solution of sugar or salt, and observe by the taste of the liquid that the solution is unaltered by passing through the paper.

Substances insoluble in water.—It has been found that salt, washing soda, nitre, etc., are soluble in water, many things, on the contrary, will not dissolve in water, and these are spoken of as insoluble. Sand, gravel, coal, camphor, are all instances of substances insoluble in water.

But though camphor will not dissolve in water, yet it disappears when shaken up in spirits of wine. And as camphor is soluble in spirits of wine, a solution of camphor in spirits of wine can be made. Shellac is another substance which will dissolve in spirits of wine and not in water; such a solution, in fact, makes a varnish used for covering some kinds of furniture. Sulphur, again, though it will not dissolve in water, and only very little in spirits of wine, disappears very quickly if placed in the nauseous liquid called carbon bisulphide.

Substances held in suspension in water.—Substances which are insoluble in water will, if finely powdered and stirred up with water, often take a long time to settle; that is, the fine particles remain *suspended* in the water for a long time. The lighter the particles are, the longer the time it

takes for them to settle and for the water to become clear. If, on a rainy day, a glassful of muddy water is taken from the gutter, and then placed on one side, the particles can be watched settling to the bottom. Those substances which, like the mud, are spread throughout the water **without being dissolved** in it are said to be *held in suspension*. The rate at which these suspended particles settle to the bottom to form a *sediment* depends upon their density and their size, large particles settling more quickly than small ones of equal density, and conversely. Whilst of particles of equal size, light particles take a long time, heavy particles only a short time to sink.

Suspended substances can be removed by filtering.—It is easy to separate suspended impurities from water. The process by which this is done is called **filtration**. Many substances are used through which to filter water containing particles in suspension. Chemists most commonly use paper which has not been glazed; such paper has been already found to be porous (p. 10); although the holes through it are large enough to let water pass, yet they are not large enough to let the suspended substances go through. In consequence, these particles are left on the paper in the funnel, and the water which trickles through is quite clear. It must be remembered, however, that it is impossible to get rid of substances *in solution* by filtering the liquid. Dissolved material passes through the holes in the paper with the liquid in which it is held in solution.

Other substances besides unglazed paper are used sometimes in filtering. Thus, often the water supply of a town is filtered through beds of sand. Household filters are made with pieces of charcoal for the water to trickle through, and in some others a particular kind of porous iron, or porcelain, is employed. Every filter requires to be cleaned frequently, or it gets clogged with impurities from the water which has filtered through it.

21. SOLUTION OF LIQUIDS.

i. **Mixture of alcohol and water.**—Pour some water into a bottle and then some alcohol,[1] and shake them up together. Observe that the alcohol disappears in the water or dissolves in it.

ii. **Solution of ether in water.**—Shake up a little water and ether. Observe the separation into two layers ; ether above and water below. Pour the mixture into a funnel the bottom of which is closed by your finger. Allow a little of the lower liquid to run into a test-tube. Insert into the test-tube a cork fitted with a glass tube. Boil the liquid and apply a light to the tube, and observe that an inflammable gas is at first evolved (Fig. 16), and after a while only steam passes off. Add a little of the upper layer to some white anhydrous blue vitriol, prepared as in Expt. 11, ii., and observe that it is turned blue.

FIG. 16.—Solubility of ether in water.

iii. **Liquids insoluble in one another.**—Shake up together some olive oil and water, and allow the mixture to stand for a short time. Notice that the liquids separate into two layers, the lighter being on the top. Which is the lighter? Boil away some of the lower water layer, and notice there is no evidence of contained oil.

iv. **Gas dissolved in water.**—Examine a bottle of soda-water. Notice that it appears clear and bright, and seems to have nothing dissolved in it. Uncork, or otherwise open it. Bubbles of gas escape. A lighted taper held to the mouth of the bottle has its flame put out by the gas which is given off.

Miscible liquids.—When alcohol and water are mixed together it is seen that one liquid alone results, and it would be obviously equally correct to call this a solution of alcohol in water or a solution of water in alcohol. Glycerine and water form a similar couple, they always form a single liquid no matter in what proportion they are
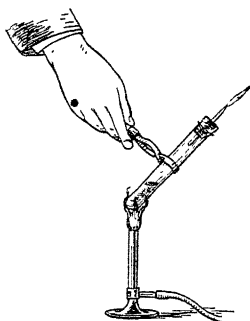
[1] Ordinary methylated spirit will not do, as it forms a milkiness with water. If pure spirits of wine cannot be obtained, whisky or brandy will do.

added together. We may say they are mutually soluble, or dissolve one another, *in all proportions*. Such liquids are said to be **miscible** or to *mix* with one another.

Liquids soluble but not miscible.—In the case of ether and water, however, a different effect is obtained. The water and ether do not form a single, or homogeneous, layer, but separate into two distinct layers. Of these it might be at first thought that the lower was pure water and the upper pure ether. It is seen, however, that the lower layer contains ether *dissolved* in the water, and it may also be proved that the upper layer, consisting *chiefly* of ether, contains also dissolved water. This may be done by adding white anhydrous " blue vitriol," which becomes blue when any water is added to it. These two liquids are, hence, mutually soluble, but not in all proportions. They *do not mix* with one another. A liquid may, hence, be soluble in water, or other liquid, without being able to mix with it. It is evident that a liquid cannot mix with water without being soluble, for miscibility is the same thing as solubility in all proportions. A solution of a liquid in water, or its mixture with water, is frequently spoken of as the liquid diluted; thus, a mixture of alcohol and water is termed dilute alcohol, etc.

Liquids insoluble in one another.—If, however, oil, water, and mercury are shaken up together, and then left to stand for a time, they will be found to separate from one another and lie in different layers—the mercury at the bottom, oil at the top, and water between the two (Fig. 17), each of these layers containing none of either of the others ; the water containing no oil or mercury, etc. Here, then, we have examples of liquids which do not dissolve in one another and do not mix. Oil and water will not mix nor dissolve in one another, quicksilver and water will not mix nor dissolve in one another, and so on.

Some gases dissolve in liquids.—When a bottle of soda-

water is opened, bubbles of gas rise out of it.   The gas has
evidently been dissolved in the
liquid.   This is only one of
many instances of gases which
will dissolve in liquids.   There
is a large amount of this gas dis-
solved in soda-water.

**Importance of air dissolved in
water.**—Rain in falling through
the air dissolves some of it in
its passage to the earth.   The air
thus dissolved in the water serves
an important purpose.   Both
animals and plants must have
air to breathe.   As is well known,

FIG. 17 —Separation of oil, water,
and mercury into three layers.

some animals and plants live in water, and these depend
upon the air which is dissolved in the water.   When water
is boiled, the dissolved air which it contains is driven out of
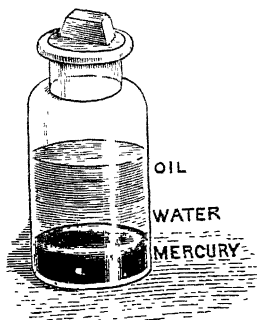it by the heat.

22. INDESTRUCTIBILITY OF MATTER DURING SOLUTION.

i. **No change of weight by solution.**—(*a*) Put some warm water
in a flask, and some salt on a piece of paper.   Counterpoise the
flask of water and the paper of salt together, and then dissolve
the salt in the water.   The total weight remains unaltered.

(*b*) Find the weight of a flask of water.   Now weigh several
lumps of loaf sugar, and put this known weight of sugar into the
water.   When the sugar has all dissolved, weigh again.   Notice
that the flask and solution of sugar together have a weight
equal to those of the flask of water and sugar added together.

(*c*) Weigh out a quantity of salt in an evaporating basin, and
add water until all is dissolved.   Heat gently by means of the
steam rising from boiling water as shown in Fig. 18.   Note that
a white solid remains, and when perfectly dry again weigh.
Observe that the weight is equal to the original weight of the
basin and salt.   Taste the solid and observe it is unchanged
salt.

No loss of material during solution.—When a substance

simply dissolves in a liquid, and so disappears from sight, it almost seems as if it is lost altogether. But this is not the
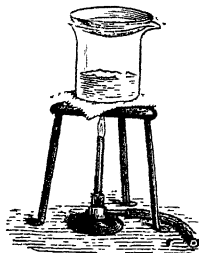

case. There is no loss whatever. In each of the experiments performed it was found that the weight of the solution was exactly equal to the weight of the solvent and the weight of the dissolved solid. It is very important to remember that there is no change of weight when a substance dissolves in water. Though sugar, for example, when dissolved in tea or water, disappears from sight,

FIG. 18.—Heating a basin by steam.

it is still in the tea or water as sugar, and it can be obtained from the water by proper means. The same fact is true of salt and other soluble solids.

23. SATURATED SOLUTIONS.

i. **Formation of saturated solution.**—Procure a supply of alum (or nitre) and powder it. Put some of the powered solid into a flask and add water. Shake them up together for some time, and if all the powder dissolves add more and shake again. Continue this addition of the powder and the shaking until some powder remains undissolved, however much it is shaken. A *cold* saturated solution, that is, a solution containing as much of the solid as it will hold, is procured.

Now warm the cold saturated solution. The powder which before remained at the bottom of the flask dissolves. Continue to add more alum, or nitre, and notice that a great deal must be added before you obtain a *hot* saturated solution.

ii. **Formation of crystals.**—Place the hot saturated solution on one side to cool. As cooling proceeds, some of the alum, or nitre, separates out in clear, well-formed crystals, because as the solution cools it cannot dissolve so much alum as before.

If time permit, make saturated solutions, at the laboratory temperature, of salt, washing-soda, borax, and lime.

iii. **Relative solubilities.**— Find the weight of a small evaporating basin. Into it, by means of a pipette, put 10 cubic centimetres of the saturated solution of alum or nitre previously made. Place the basin upon a sand-bath and gently heat the liquid, which will

gradually disappear, leaving the solid behind in the basin. If the liquid 'spits' when almost dry, use the method employed in Expt. 22, (c). When the residue is quite dry, allow the basin to cool, and weigh it again. The increase in weight gives the amount of alum or nitre dissolved in 10 cubic centimetres of the saturated solution.

Find out, in the same way, the amount of solid in 10 cubic centimetres of each saturated solution prepared.

Saturated solutions.—When any given amount of water has dissolved as much of a solid as it can be made to, without warming or assisting it in any other way, it is said to be saturated. But though cold water, for instance, may be saturated with any particular solid, such as sugar, it can, if warmed, be made to dissolve more sugar. Though there are some exceptions, it may be regarded as the general rule that water and other liquids will dissolve more of a solid when they are warm than when they are cold. In most cases the amount of solid which will dissolve goes on increasing as the temperature of the water is increased. In general, therefore, the cooler the water the less of a solid will it dissolve. Now suppose warm water is given as much sugar, salt, alum, or any substance of this kind as it will hold, and is then cooled, what happens? It has to give up some of the substance, for it cannot hold as much as when it was warm. When tea has been made very sweet, some of the sugar is left on the bottom of the cup as the tea cools. This is because, though the tea was able to dissolve a certain amount of sugar when hot, it could not hold so much when cold, and therefore a little of it was deposited upon the bottom of the cup.

24. MEASUREMENT OF SOLUBILITY.

i. **Determination of solubility.**—(a) In a large flask place a considerable quantity of nitre and water, and, having placed a thermometer in the flask, warm until it reaches a temperature of say 70°. Endeavour to keep the temperature constant at about this value for a considerable time, shaking and stirring the flask well throughout. If all the nitre dissolves, more must be added

and the flask kept at a constant temperature with constant stirring or shaking for at least 15 minutes after the last addition. Allow the solid to settle, and, by means of a pipette which has been just heated to 100° by boiling water,[1] draw off 10 or 25 c.c. of the liquid and transfer it to a weighed evaporating basin. If this is done quickly no solid will separate out and remain in the pipette. Weigh the evaporating basin with the solution.

Repeat with different evaporating basins, each previously weighed, at temperatures of about 60°, 50°, 40°, 30, and 20 respectively, allowing the solution to cool to about the desired value, at which it should be kept approximately constant by means of a small flame.
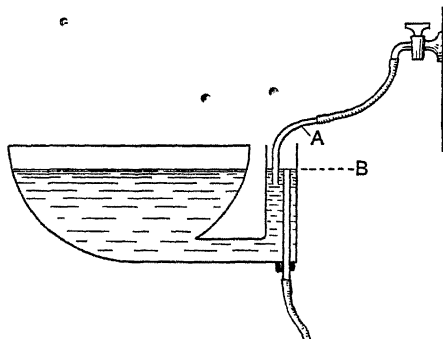

FIG. 19.—Water-bath.

If ice is available, cool a small beaker of the solution by immersing it in a basin of ice until the temperature is about 1°, and then, as before, transfer 10 or 25 c.c. to a weighed evaporating basin.

If no water-bath [2] (Fig. 19) is available obtain a large saucepan and make for it a cover of tinplate or copper with circular holes in which the evaporating basins will fit, but will not drop through. Place water in the saucepan and the evaporating basins on the cover. Boil the water in the saucepan and it will be found that the water in the evaporating basins will gradually pass away

[1] As the pipette will be hot some india-rubber tubing on the end is advisable.

[2] A water-bath is convenient, as it may be left for some time without becoming dry; the water enters by the tube A (Fig. 19) and runs off by the tube B, so that the level always remains at the level of the top of B.

without, however, any *spitting* of the solid, which, as you will have found in Expt. 23, iii., is liable to occur.

In this way evaporate 10 c.c. of each of the solutions prepared and again weigh the evaporating basins. From your results determine, for each temperature, the solubility of the nitre in grams dissolved per litre of water. Record your results thus :

Temperature = 60° C.

Weight of evaporating basin and solid from
$$10 \text{ c.c. of solution} = 43\cdot62 \text{ gms.}$$
Weight of evaporating basin = $36\cdot67$ gms.

Weight of solid dissolved in 10 c.c. of solution = $6\cdot95$ gms.

10 c.c. contained $6\cdot95$ gms. dissolved nitre.
1000 c.c. contained 695 gms. dissolved nitre.
Solubility at 60°, - - - 695 gms. per litre.

Calculate also the solubility in grams, per 100 grams of water as below :

Temperature, 40·5° C.

Weight of evaporating basin and solution,	-	42·71 gms.
Weight of „ „ „ solid, -	-	35·41 gms.
Weight of water, - -	-	7·30 gms.
Weight of evaporating basin and solid, -	-	35·41 gms.
Weight of evaporating basin, - - -	-	30·96 gms.
Weight of solid, - -	-	4·45 gms.

Quantity dissolved by 7·3 gms. of water = 4·45 gms.
$$\text{„}\qquad\text{„}\qquad\text{„ 100 „}\qquad\text{„}\qquad =\frac{445}{7\cdot3}\text{ gms.}$$
$$=61 \text{ gms.}$$

Solubility at 40·5° is 61 grams per 100 grams water.

(*b*) Repeat your experiments, using others such as salt, washing-soda, etc., and thus obtain the solubility for each in grams per litre, and in grams per 100 grams of water, at various temperatures.

ii. **Comparison of solubility curves.**—Construct, in the manner shown on p. 45, on the same piece of squared paper, the solubility curves, in grams per litre, for salt and for nitre, and compare their appearance.

iii. Construct a similar curve to represent the solubility in grams per 100 grams of water for the two substances.

Effect of temperature on solubility.—It will be clear from your experiments that not only does the temperature of the water affect the solubility, but that it affects different substances to a considerably different extent. Thus, whereas in the case of the nitre the quantity dissolved increased greatly as the temperature is raised, yet in the case of common salt the quantity dissolved in the cold water is almost as great as that dissolved in the hot water. In the first case the solubility increases considerably with the temperature ; in the second case the solubility varies but slightly with temperature. As a general rule it may be stated that solids are more soluble in hot than in cold water, or that the solubility of solids increases with the rise of temperature. There are two common methods of expressing the solubility of a solid :

(1) By stating the weight in grams dissolved in 1 litre of the solution.

(2) By stating the weight dissolved by 100 grams of pure water.

Graphic representation of solubilities—Solubility curves. —The solubility of substances at different temperatures may be readily and conveniently shown in a graphic manner. It has, for example, been found that one litre of nitre solution contains 740 gms. at 70°, etc. Now, on squared paper we may rule two lines at right angles, such as OX, OY (Figs. 20 and 21), and we may mark temperatures along the line OY letting each division indicate say 2 degrees. So that, if the temperature 0° be marked at O, then 10° is marked at the fifth line, 20° at the tenth, etc. Along the vertical line, OX, we mark the weight of nitre contained in each litre of solution, allowing each division to represent say 20 gms. Then, if at 10°, the weight dissolved in 1 litre of solution is 150 gms., we mark the point (a) which represents the temperature 10° (being 5 divisions from the left hand line OX), and also represents 150 gms., being 7·5 divisions above OY.

In a similar manner, we find the position of the other points on the solubility curve, and the complete curve can then be constructed by drawing a continuous line to pass through these points.
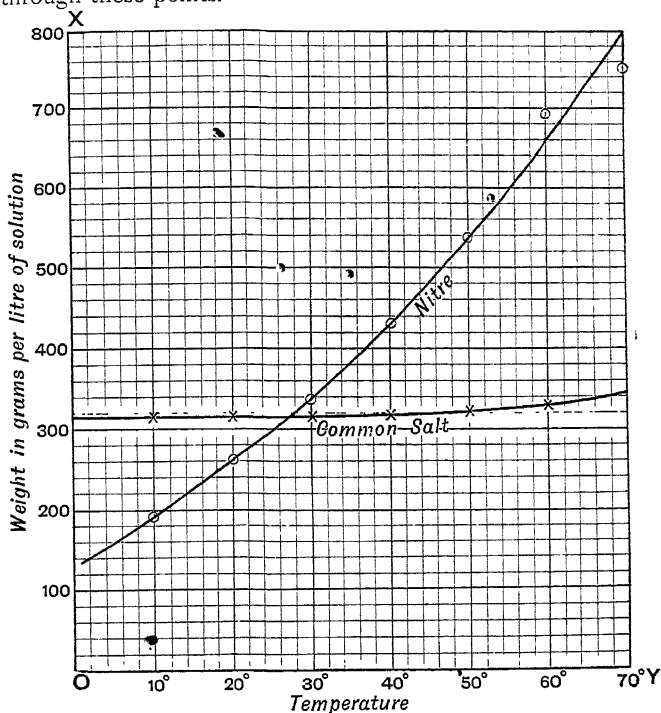

FIG. 20.—Solubility curves of common salt and nitre.

Solubility curves of nitre and salt.—After having drawn the curves for nitre and salt, as described in Expt. 24, ii., you will notice a great difference between these two curves. Whereas in the case of nitre the curve rises rapidly, in that of the salt it rises but slowly.   That is, in the case of nitre the solubility increases rapidly with the temperature, in the

case of salt it increases but slightly. Nitre is much more soluble in hot water than in cold. Salt only slightly more so.
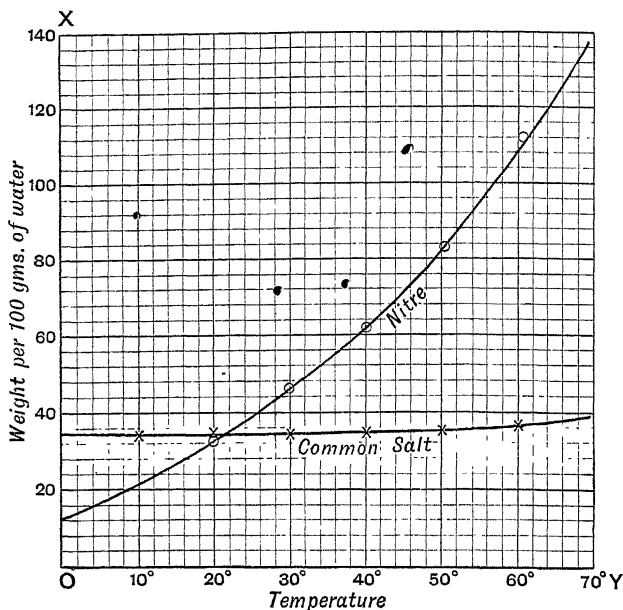

FIG. 21 —Solubility curves of common salt and nitre.

25. CRYSTALLISATION.

Production of crystals.—(*a*) Make a saturated solution of soda at the ordinary temperature of the laboratory and place it aside in an evaporating basin, so that the water slowly evaporates. Examine after a few days and note the crystals formed.

(*b*) Evaporate a solution of soda by means of a Bunsen burner, and observe that a white powder is formed and no crystals are evident to the eye.

(*c*) Make a warm saturated solution of soda, or alum, in a

flask and place it aside so that it may slowly cool. Observe after a day has elapsed that fine large crystals are present.

(*d*) Make a similar saturated solution, and having poured some into an evaporating basin, float the latter on cold water in a trough so that it shall cool more rapidly. Observe that the crystals formed are much smaller than in (*c*).

(*e*) Melt some powdered sulphur in an iron spoon or cup, and then allow it to cool slowly. When a solid crust has been formed over the top, make two or three holes in it, and pour off the remaining liquid sulphur. When the sulphur is cool, examine the inside of the spoon or cup, and notice the fine needle-like crystals of sulphur.

(*f*) Having melted some sulphur as above, pour it when melted into some cold water. Examine the product formed, and observe it has still a crystalline appearance; the crystals are so small that individual crystals cannot be distinguished.

The formation of crystals.—During the examination of saturated solutions it is found that when a solid separates out from its saturated solution it frequently does so in the crystalline form. Many crystals have been examined, and it has been found that in many cases the crystals contain water, called water of crystallisation. In other cases, however, the crystals are anhydrous containing no water, *e.g.* salt. A solid may be made to separate from the saturated solution in two ways: (1) by the evaporation of the water; (2) by cooling the solution. In both cases the process may be effected quickly or slowly, thus the water may be allowed to evaporate slowly into the air without the application of heat, or the evaporation may be hastened by heating. It is found that the crystals formed in the first case are much larger than those obtained in the second case. This is also true when the crystals are formed by cooling a saturated solution. If the cooling is rapid the crystals formed are small; when the cooling is slow, however, fine large crystals may be obtained. Crystals are also produced without the intervention of a solvent. When a substance changes from the liquid state to the solid it frequently forms crystals. In this case also the crystals are larger and more clearly defined

when the solidification proceeds slowly than when it proceeds
rapidly. Sulphur, for example, when it slowly solidifies forms
fine, glassy crystals. Many of the crystals in rocks have
probably been thus produced.

26. EVAPORATION.

i. **Evaporation of water.**—Put a *small* quantity of water in an
evaporating basin and leave it undisturbed on a shelf of the
laboratory. Examine after a few days, and notice that the water
has disappeared and that the basin is dry. The water is said to
have *evaporated*.

ii. **Evaporation assisted by heat.**—Put some water in an eva-
porating basin over a water bath as described in Expt. 24, i. ;
or, over a beaker of boiling water as in 22 (*c*). Heat by a
Bunsen burner so that the water in the water bath, or beaker, is
kept gently boiling. Observe that although there is no bubbling,
or boiling, of the water in the evaporating basin, yet it gradually
disappears or evaporates ; and in a short time, perhaps an hour,
the basin is perfectly dry.

iii. **Evaporation and boiling.**—Examine carefully the water
boiling in a beaker or in a flask, and observe that, when boiling
begins, bubbles of steam are formed everywhere throughout the
liquid, and that the bubbles rise and burst at the top.

iv. **Cooling produced by evaporation.**—(*a*) Sprinkle a few drops of
(1) spirits of wine, (2) carbon bisulphide, (3) ether on your hand
in succession. Notice that the liquid soon disappears, and its
presence in the air can be detected by its smell. The rate at
which the liquid evaporates is increased by waving the hand
about, or by directing on to it the stream of air from a bellows.

(*b*) Place some ether in a small beaker and cause it to
evaporate *rapidly* by blowing a current of air through it.
Observe the great fall in temperature. Place in the beaker a
small test-tube containing a *little* water, and observe that
the water becomes frozen.

Evaporation.—If a saucer of water is left for a few days,
the water will disappear, or, as is generally said, dry up.
The water can be made to disappear more quickly by gently
heating it. When a solution containing salt, or sugar, is made
to dry up in this way, the salt, or sugar, does not disappear,
but remains in the saucer. The name given to this process

of turning a liquid into a vapour is *evaporation*.   The solid left behind is spoken of as a *residue*.

Everyday examples of evaporation.—In summer it is not long before a road becomes dry again after having been well watered by a water-cart.   Wet clothes hung upon a line and exposed to the sun and air soon become quite dry.   It is a common practice to expose shallow vessels of water in rooms which are warmed during winter by coke- or gas-stoves.   From time to time such vessels have to be refilled, for the water quickly disappears and spreads, in an invisible form, throughout the air in the room.   Nearly everybody has observed the same process taking place when sitting still after some violent exertion.   The perspiration, which after such exercise stands in drops over the face and body, soon disappears by passing into the air as vapour, and leaves the skin dry.   The cooling of the skin of the body which is also noticed at the same time will have to be explained a little later.

Rate of evaporation.—It has been seen that evaporation is increased by means of a current of air, and this is a fact of general knowledge.   Clothes, for example, hung out to dry, do so more rapidly during a slight breeze; whilst a common method of drying glass articles, such as flasks, in a laboratory is to warm them and then to blow a current of air through them by means of bellows.   The reason for this is that water, or other liquid, will not evaporate indefinitely into a closed space of air; thus, if some water is placed in a closed flask, a little will evaporate, but not all.   This is because after a certain amount has evaporated the air becomes *saturated* with the water vapour, just as water becomes saturated with salt which is dissolved in it.   Thus, for a solution of salt, or nitre, for every temperature a certain concentration of salt can be reached at which the solution is saturated, and no more salt can pass into the solution, the concentration required for saturation increasing as the

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temperature rises. It is very much the same with the water vapour and, the air. When the concentration of the water vapour in the air has reached a certain amount the air is saturated, and no more water vapour can pass into it; this concentration also increases with rise of temperature.

Now, although water exposed to the atmosphere in a saucer, etc., is not in a closed space and would be insufficient to saturate the whole atmosphere, yet, if the air is still and undisturbed, the air near the water may get saturated, and it is only slowly replaced by unsaturated air. In a current of air, however, the air saturated with moisture is being continually replaced by the unsaturated air, and hence the water evaporates much more rapidly.

Cooling during evaporation.—You have already observed that a considerable amount of heat has to be given to a liquid before it can be converted into vapour. This amount of heat is termed the *latent heat of vaporisation*, and you have measured it in your physical work. This absorption of heat is true whether the liquid boils, or whether it merely evaporates more or less slowly. When a liquid boils the necessary heat is supplied by the laboratory burner or other source of heat. But when liquid left in a saucer, or put on the hands, evaporates without an external source of heat being employed, it has to obtain the necessary quantity of heat from the remaining bulk of the liquid, and from the objects with which it is in contact. If the evaporation is very slow the heat is only abstracted slowly, and the objects, or liquid, in contact with warmer air, are able to take heat from the air, so that their temperature does not fall to a marked extent. When the evaporation, however, is rapid, then the heat abstracted from the liquid, for the evaporation, cannot be sufficiently rapidly compensated by gain from the air, etc., and the temperature falls very considerably.

Differences between quiet evaporation and boiling.—It is very instructive to watch some vigorously boiling water

in a flask or beaker, and to compare it with the evapora-
tion of a solution in an evaporating basin, where only a
small flame, or no flame at all, is used.   In the case of the
solution being gently evaporated all the vapour is formed *at
the surface* of the liquid, and the process goes on quietly
until no liquid is left in the basin.

In the case of vigorously boiling water, bubbles of vapour
are formed *everywhere throughout the mass of the liquid.*   The
bubbles can be seen at the bottom and the sides, and they
rise from every point to the surface, each, as it escapes there,
making a little noise.   The sounds of the bursting bubbles
added together make up the " singing " or "rattling " which
is heard when water is briskly boiling in a flask or other vessel.

27. CONDENSATION AND DISTILLATION.

i. **Condensation of water vapour from the air.**—(*a*) Put some
pieces of ice into a test-tube, or other glass vessel, which is
clean and dry on the outside.   In a very few minutes the outer
surface of the vessel will become covered with moisture, owing to
the *condensation* of water vapour from the air.

(*b*) If you cannot procure ice, put a little ether in a test-tube and
make it evaporate quickly by blowing vigorously down a narrow
tube through the ether.   As the air is cooled by the evaporation
of the ether, the moisture in the air is given up and deposited on
the outside of the test-tube in the form of minute drops ; in other
words, the water vapour in the air is condensed.

ii. **Distillation of water**—Obtain a glass retort and put in some
ordinary tap-water coloured with ink, in which some sand has
been placed, and put it on a sand-bath, or piece of iron gauze ;
attach to the retort a condenser as shown in Fig. 22, and care-
fully notice the construction of this condenser.   Observe that it
consists of an inner tube through which the vapour passes,
surrounded by an outer tube through which cold water flows,
passing in at the lower end and out at the upper end.   Let the
end of the condenser tube dip into a flask.   By means of a
Bunsen burner, boil the water in the retort and keep the con-
denser cool by the stream of cold water.   Notice that the steam
which passes over into the flask is condensed again into water,
which is quite clear and tasteless, both dissolved and suspended
matter remaining in the retort.

Condensation.—By heating, a liquid is changed into vapour. This change may take place slowly and gently, as in evaporation; or quickly and vigorously, as in boiling. But by whichever process vapour is obtained, it can, by the
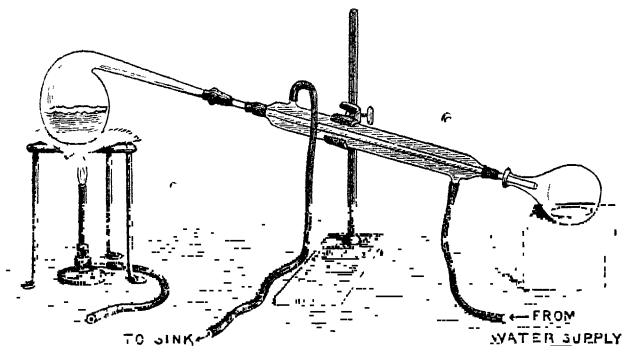

FIG. 22.—Distillation of water.

reverse plan of cooling it, be reconverted into liquid. This change, from the state of vapour back again to the condition of liquid, is called *condensation*; the vapour is condensed to a liquid. Thus, a cold plate, held close to the spout of a kettle from which steam is coming, will cool the steam, condensing it into water, which will be seen trickling down the plate.

Most people have noticed the condensation of vapour taking place at some time or other on a cold day. For instance, if the doors and windows of a room are kept tightly closed, and there is a good fire burning, the inside of the window panes soon becomes covered with moisture, which, forming drops, trickles down the glass and collects on the window frame as liquid water. The water must evidently come from the air in the room.

The air outside a room is, in winter, much colder than that inside. This cools the glass of the windows very much,

and consequently the air next to the cold surface itself becomes cooled, and then cannot hold so much vapour as when it is warm, and some of the vapour which can no longer be held by the air is changed into water. So that condensation is caused when air containing water vapour is cooled.

Distillation.—The change of liquids into vapours by heating them, and the condensation of the vapour into liquid by cooling it, is employed in an important process called *distillation*. This plan is frequently made use of for purifying water and other liquids. Perhaps the most useful application of distillation is to obtain fresh water for drinking purposes from sea-water, or other water not fit to drink. Large ships, carrying as they often do more than a thousand people, cannot take enough fresh water on board for the needs, throughout a long voyage, of so many persons. Instead of attempting this difficult task, it is the custom to change sea-water into fresh water by distillation; as the water boils it is converted into steam, and this steam is condensed again into water in the cooled flask. If some ink and salt were first added to water in a retort, they would be left behind on distilling, and the water found in the flask would neither taste of the salt nor be coloured by the ink; it would be purified from these by distillation. Water can, by distilling in this manner, be purified from very many impurities, such as dissolved solid matter, the condensed vapour being pure water.

Distilling apparatus.—A form of *still* used for obtaining large quantities of distilled water is shown in Fig. 23. In principle it is essentially similar to the retort and condenser employed in Expt. 27, ii. The water is boiled in a large copper vessel *B*, and the vapour passes from this vessel through a spiral tube *D*, which is kept cold by a stream of cold water. The condensed water drops from the end into a vessel placed to receive it.

To obtain *absolutely* pure water, however, is a matter of considerable difficulty, since, although dissolved solids are got rid of, yet volatile substances which will evaporate away with the water and condense with it may still be present, as
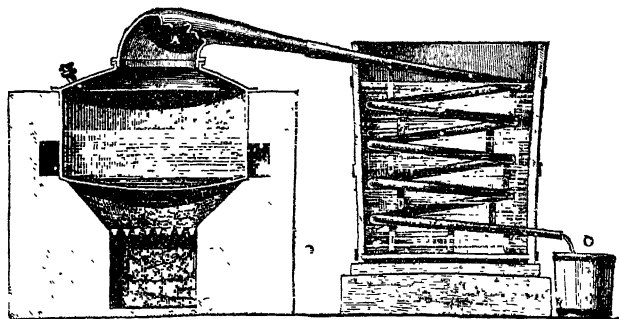

FIG. 23.—Still for distillation of water.

may also dissolved gases and substances dissolved from the walls of the retort or condenser, since few substances are absolutely insoluble in water. However, *absolutely pure* water is very seldom required by the chemist or anyone else, and for almost all purposes the ordinary distilled water is as pure as is required.

QUESTIONS ON CHAP. III.

1. What do you mean by a soluble thing? Give examples.
2. Explain how you would proceed to make a solution of table salt.
3. How would you obtain clear water from muddy water?
4. What do you mean by substances being held in suspension in water?
5. What kind of impurities cannot be got rid of by filtration?
6. Explain the terms—soluble, insoluble, filtration, and 'held in suspension.'
7. What is really meant when we say two liquids will not mix?
8. Give instances of (1) liquids which mix, (2) liquids which do not mix.
9. Will gases dissolve in water? Give reasons for your answer.
10. What would you notice if you were to shake (1) some ink and water together, (2) some oil and water?

11. How would you prove that there is no loss of matter during solution?

12. How would you obtain salt from a solution of salt in water?

13. What do you know about evaporation?

14. Describe some practical applications of the processes of solution and evaporation.

15. What is a saturated solution? Describe how to make one.

16. Which will dissolve more sugar, warm or cold water? What is the general rule about the effect of an increase of temperature on the dissolving power of water?

17. What do you know about water as a solvent?

18. Describe fully what happens if you gradually cool a hot saturated solution of alum.

19. How would you proceed to show that: (*a*) Rain water is free from dissolved solids. (*b*) Sea water contains very much dissolved matter.

20. A ship which floats up to a certain mark in river-water is found to float more out of water when at sea. How can you account for this?

21. Spring and river water are not pure. Describe carefully how you would prove this, and say how you would estimate the quantity of dissolved matter in a gallon of each.

22. Distinguish between soluble and insoluble substances. Which of the following are soluble in water? Sand, sugar, table-salt, camphor, sealing-wax, baking-powder, and coal.

23. Some solids which do not dissolve in water are soluble in other liquids. Give all the instances of this with which you are acquainted.

24. What class of impurities can be removed from water by filtering? Explain clearly how you would proceed to get clear water from muddy water.

25. Describe fully how pure water can be obtained from sea-water.

26. Do gases dissolve in water? If you think they do, name two or three soluble gases.

27. How is it that clothes placed on a line to dry are ready to be taken in sooner on some days than others? What sort of day does the laundress consider a good drying day?

28. In both quiet evaporation and boiling, a liquid is converted into vapour. What differences are there between the two processes?

29. What is meant by condensation? Name two or three familiar instances of the condensation of water vapour.

30. A mixture of salt and powdered glass is given to you. How could you (*a*) obtain separately the two constituents, (*b*) find the quantity of each in the mixture? Could you separate sugar from salt in the same way.

31. What do you mean by a saturated solution? How would you prepare a saturated solution at a given temperature? What is generally the effect of cooling a saturated solution?

32. A white powder is shaken up with water. How would you ascertain whether any of it dissolves?

33. What is meant by the statement that two liquids mix. Give examples. Can a liquid be soluble in another liquid without being capable of mixing? If so, give an example.

34. How may fresh water suitable for drinking be obtained from sea water?

35. How would you endeavour to obtain large crystals of alum?

36. Describe an experiment to show the solubility of ether in water.

37. Sand and salt are stirred up in a bottle containing water in which some gas is dissolved. What methods would you use to recover the sand, salt, and gas from the water?

38. Give examples of the solvent power of water. How could you determine whether a sample of water contained dissolved solid matter?

39. Describe how to make soda crystals from powdered carbonate of soda.

40. If you were given some powdered alum, explain how you would proceed to make a crystal of alum.

41. Describe the method you would adopt in order to determine the influence of temperature on the solubility of a substance.

42. Does the solubility of all substances vary equally for the same change of temperature? State the experimental evidence on which you base your answer.

43. From the following table plot a solubility curve, and use it to determine the weight of substance contained in 100 grammes of solution saturated at 60° C.

Weight of substance dissolved in 100 grammes of water.							Temperature
2·8 grammes,	-	-	-	-	-	-	0° C.
4·7 ,,	-	-	-	-	-	-	10° C.
7·9 ,,	-	-	-	-	-	-	20° C.
14·5 ,,	-	-	-	-	-	-	35° C.
24 ,,	-	-	-	-	-	-	48° C.
58 ,,	-	-	-	-	-	-	70° C.
96 ,,	-	-	-	-	-	-	86° C.
201·5 ,,	-	-	-	-	-	-	100° C.

CHAPTER IV.

ACTION OF HEAT ON METALS.

28. RUSTS.

i. **Examination of rusts.**—(a) Prepare some rust of each of the metals examined in Chap. ii., and carefully examine the rusts according to the scheme in Chap. ii. Note that the rusts differ in colour, but resemble one another in the following points: They are insoluble in water. They are not altered by heating, except to this extent, that the rusts of zinc and tin turn yellow when hot, returning to the white state when cold; whilst the rust of lead melts and darkens in colour, but reverts to its original form when cold. Compare these properties with those of the original metals.

(b) Tabulate the properties of these rusts thus:

Rust of	Colour.	Action of water	Action of heat
Copper, -	Black	None	None.
Iron, - -	Reddish-brown	,,	,,
Lead, - -	Yellow	,,	Melts and darkens.
Zinc, - -	White	,,	Becomes yellow.
Tin, - -	,,	,,	,,
Magnesium,	,,	,,	None.

It has been already noticed that a number of the metals which have been examined undergo certain changes when heated. For example, copper, lead, iron, zinc, tin, magnesium, have all been found to lose their bright, metallic lustre when heated and to become coated with a tarnish which exists only on the surface, for if the tarnish is scraped off the bright metal is again seen below. In the case of

magnesium, the metal, if heated sufficiently, burns forming a white ash. The tarnish which forms on iron is termed *iron rust*, and we may also similarly call the tarnish in other cases *rust*; thus, we may speak of the rust of copper, rust of tin, etc.

29. CHANGES OF WEIGHT DURING RUSTING.

It is now necessary to examine further the general case of rusting.

i. **Iron in rusting gains in weight.**—Weigh carefully a watch glass with some iron filings or tacks ; add a few drops of water (because iron rusts best in the damp) and allow it to stand. At the end of a few days warm gently to drive off the water, and when *quite dry* again weigh, and note carefully the weight. Has the iron gained or lost in weight?

ii. **Copper in rusting gains in weight.**—Weigh carefully a crucible containing some piece of thin, bright copper foil. Place it in a furnace for some hours until the copper has been well covered with its black rust. Allow it to cool and again weigh. Compare with the previous weight, and notice that, like iron, copper gains in weight by rusting.

iii. **Density of iron rust.**—Determine the density of some iron and of some iron rust, using a density bottle, such as you have used for the determination of the density of powders or small objects. Compare the values.

The nature of rusting.—In order to determine the nature of the change which occurs during the rusting of the metal, the first question to be considered is, what takes place during the rusting? Does the metal, for example, lose or give up anything? Or, on the contrary, does it gain anything? To answer these questions it is only necessary to weigh carefully a watch glass with some iron filings or tacks, and allow the iron to rust in the air, which will take a few days; on warming the rusty filings until quite dry and weighing again, it will be found, if care has been taken, that the iron has increased in weight, and the experiment has furnished the necessary answer, so that we may write :

Iron gains in weight during rusting.

Similarly copper when heated so as to form its black rust is also found to gain in weight, so that the rust is heavier than the original copper.

Copper gains in weight during rusting.

It is not so easy directly to determine quantitatively whether the other metals have gained in weight, but the experiments can be performed and analogous results are found. It may in fact be stated as a general law that

Metals gain in weight during rusting.

For further investigation, the rusting of iron, which proceeds without the application of heat, will serve as a convenient example. The rust has already been examined and compared with iron. Inasmuch as it has been found that the rust has a greater weight than the iron it is interesting to find whether the increase in weight is accompanied by an increase in density.

Experiment shows that although the iron increases in weight during rusting, yet its density decreases, the density of rust being less than that of metallic iron. It is clear, therefore, that the volume must increase during the rusting.

30. CONDITIONS OF RUSTING.

i. **Moisture necessary for rusting.**—Take two evaporating basins containing iron filings ; to one add a few drops of water, but leave the other quite dry. Allow both to remain in the air for about a week and then examine both basins. Observe that the iron in the basin to which water was added has rusted to a much greater extent than that in the dry basin.

ii. **Air necessary for rusting.**—Take two test-tubes and put some water in each so that they are about half full. Into one put some iron filings, but before placing the iron filings in the second, boil the water for a couple of minutes, so that all the air which may be contained dissolved in the water shall be expelled. Then place in the iron filings, boil for a few seconds longer, and then put above the water a little vaseline, which, you will observe, immediately melts and forms a layer which will prevent any air being afterwards dissolved. After some days again examine these two tubes, and note whether the iron

has rusted (1) in the tube containing water with dissolved air, (2) in the tube containing water quite free from dissolved air.

Necessity of moisture and air for rusting.—When iron is moistened and exposed to the air it rusts and at the same time gains in weight. It is important to find whether the moisture and the air are both necessary, and if not, to which the rusting is due. For this purpose dry and moist iron are exposed to air, and the result clearly shows that the moistened iron rusts far more readily than that left dry, and the experiment indicates the necessity for moisture if rust is to form. In order to see whether water alone is sufficient, another experiment is performed, whereby iron is first left in water containing air, and, secondly, in water free from air. It is found that although in the former case a copious layer of rust forms over the iron, yet in the latter case, the iron, even after the expiration of a few days, is still completely unrusted. This fact shows clearly that air as well as moisture is necessary, and that without air, even in the presence of plenty of water, no rusting occurs.

31. EFFECT OF RUSTING ON THE SURROUNDING AIR.

i. **Iron in rusting abstracts a portion of the air.**—(*a*) Place some iron filings in a muslin bag and tie the bag to a piece of glass rod. Moisten well (better dip it in a solution of sal-ammoniac) and place it in a bottle of air inverted over water (Fig. 24). Examine after a few days. It will be seen that the water has risen in the glass, showing that some part of the atmosphere has been abstracted by the iron in rusting.

(*b*) Tightly place your hand, or a card, under the mouth of the jar so as to allow no water to escape, set the jar upright and place a burning taper into it. Place also into the jar some burning sulphur or phosphorus, the substances being placed in a deflagrating spoon (Fig. 41). Note what happens, *but do not throw away the water.*

ii. **Volume of the part of the air abstracted.**—Next measure in a graduated vessel the quantity of water in the bottle. This is equal to the quantity of gas which has been used by the iron.

Also measure the quantity of water the bottle holds.   This is the quantity of air it originally held (Fig. 25).

111. **Repetition and verification of previous results.**—Repeat the previous experiments, using bottles of different sizes, and allowing the iron to remain for different lengths of time (but more than three days).   Use at least 5 bottles.   Compare the results carefully, and examine the iron left in the muslin bags.
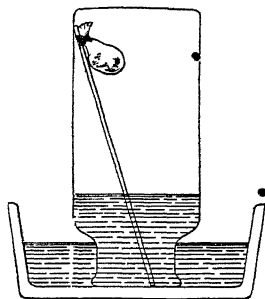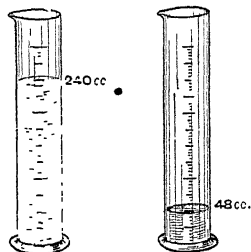

FIG. 24 —Abstraction of a portion of the air during the rusting of iron.

FIG. 25 — Proportion of air abstracted during the rusting of iron in a closed jar.

Portion of air abstracted during rusting. — We have found that iron gains in weight when it rusts, and also that the presence of air and of moisture is a necessary condition for rusting.   It is clear that the iron takes up some substance from somewhere, since it gains in weight, and the question naturally arises, does this gain come from the water which was added, or from the atmosphere?   To decide this point, the rusting of the iron is allowed to take place in the apparatus shown in Fig. 24.   The iron filings are enclosed in a muslin bag, which is placed in a bottle of air inverted over water.   On examining after a few days, the water will be seen to have risen in the bottle, showing that some part of the atmosphere has been abstracted by the iron in rusting.

Residual gas will not support combustion.—The gas which is left in the bottle, moreover, will not allow a lighted taper to burn in it, but immediately causes it to be ex-

tinguished. In the similar manner, sulphur or phosphorus will not continue to burn, so that it is clear that the gas left in the bottle *differs from air* in the respect that it does not allow substances to burn in it.

If the volume of air left after iron has rusted in a closed space is compared with the original volume, it will be found that one-fifth of the air in the bottle is abstracted by the iron in rusting.

One-fifth of the volume of air is abstracted.—These results are of such fundamental importance that they should be completely verified. If a class of twenty or more are working together, the verification may be obtained by a comparison of the different results ; if, however, a number of results are not available it is most necessary that the experiment should be repeated. It will be found that, in all cases, only about one-fifth of the air is abstracted ; whilst, further, that a quantity of iron remains unrusted, so that when one fifth of the air has been taken by the iron it cannot abstract anything more from the remaining four-fifths. These experiments thus show that one-fifth of the air is different from the other four-fifths, and we must consider the air as made up of two portions, one of which is concerned in burning, while the other (the four-fifths) does not support burning, or combustion. We may therefore state :—

Iron in rusting gains in weight, taking some material from the air, and this material is the part of the air concerned in burning.

If we call the part taken up by the iron the *active* part of the air, and the other part the *inactive* part, it is seen that the results further show that air contains 80 per cent. *inactive part*, which does not allow things to burn in it, so that the remaining 20 per cent. must be the part concerned in burning, which we have called the *active part*.

32. INACTIVE AIR.

i. **Iron does not rust in the inactive air.**—Obtain a bottle of inactive air as in Expt. 31, i.; and place in it another bag of bright, moistened iron filings tied to a glass rod as before. Allow these filings to remain for three or four days and then examine the bottle. Note, has the water risen further? have the iron filings rusted?
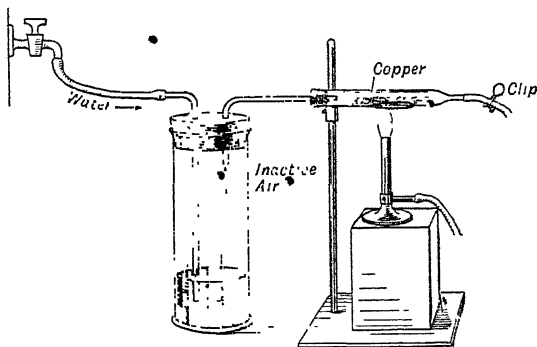

FIG. 26.—Copper does not rust in the inactive air.

ii. **Copper does not rust in the inactive air.**—Allow some iron to rust in a *large* jar or bottle of air as in Expt. 31, i., so that a considerable quantity of the inactive air is obtained. After some days place some bright copper turnings in a glass tube and connect, as shown in Fig. 26, to the jar or bottle which contains the inactive air left after the rusting of iron. This bottle must be closed by a cork [1] through which pass two tubes ; one of these, the short one, is attached by india-rubber tubing to the tube of copper, while the long tube passes to the bottom of the bottle and is attached to the water tap. When the water is turned on, the inactive gas is forced out of the bottle and fills the tube of copper. When it is considered that all the air has been displaced and only the inactive air is present in the tube, the end of the tube is closed by a pinch-cock and the tube heated. Notice that the copper retains its bright lustre and shows no trace of tarnishing. Now disconnect from the bottle, open the

[1] It is best to insert this cork, and hence close the bottle, while still inverted with its mouth below water and thereby avoid the possibility of any *air* entering the bottle.

pinch-cock, and slope the tube so that *air* may pass along it, you will observe that the copper immediately tarnishes.

Inactive air will not cause rusting.—We know that substances will not burn in the inactive part of air, and it was found also that the iron employed for obtaining inactive air is only partially rusted. Iron which has not been employed for rusting at all, and is perfectly bright, remains so in the inactive air. Further, other metals, as copper, may be heated in the inactive air without any trace of rust or tarnish, although immediately the air reaches it, it at once becomes coated with the brightly coloured film seen when copper tarnishes. The inactive part of the air therefore will support neither burning nor rusting.

33. EFFECT OF HEAT ON OTHER METALS.

i. **Action between hot copper and air.**—Place a roll of copper gauze, or some copper turnings, in a long hard glass tube provided with a bored cork at each end. Find the weight of the tube and the
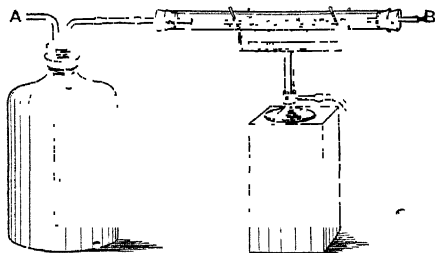
FIG. 27 —Action between hot copper and air.

copper. Connect one end of the tube with an aspirator or a bottle fitted as in Fig. 27, so that air can be made to pass through the tube. Connect the other end of the hard glass tube with a tube passing under a jar standing inverted in a basin of water. Heat the copper to redness, and drive air over it, by letting water run slowly into the corked bottle. Notice that the copper turns black, owing to its combination with the active part of the air. Allow the tube to cool, and redetermine the weight of the tube and its contents. The weight has

increased. Lift the jar from the water and insert a lighted taper into it. The taper is extinguished. thus showing that the part of air required to sustain burning has been removed.

ii. **Gain of weight during heating of magnesium.**—Weigh a crucible and lid and piece of magnesium, which folded tightly, is placed in the crucible. Heat strongly in a burner, taking care to let no fumes escape (Fig. 28). To do this, keep on the lid, and only raise it a little when the flame is removed. The magnesium is seen to burn brightly in places ; but, if care is taken, no fumes are lost. Towards the end the crucible should be strongly heated, and when the experiment is finished, the whole mass should be in the form of a white powder. Allow to cool, and weigh the crucible with the lid

FIG. 28.—Gain of weight during heating of magnesium.

and powder. Subtract the weight of the crucible and lid to find the weight of the powder. It will almost certainly be found to have increased in weight, and, if the experiment is carefully done, it will be found that the weight of the magnesium has increased by about 66.5 per cent.
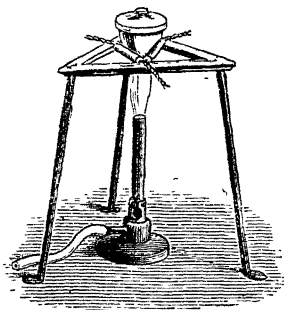
Part of the air can be removed by other heated metals. —If air be passed over red-hot copper, the copper changes in colour, and if the remaining part of the air is examined by plunging a burning taper into it, the flame is extinguished. It has been found previously also that the copper gains in weight. There is no difference between the condition of things in this case and in the rusting of iron, except that the copper must be strongly heated before it combines with the active part of air. It does so then, however, and we have already seen that copper is quite incapable of rusting in the inactive part.

Magnesium also gains in weight when heated, but in this case the magnesium actually burns and the phenomenon appears at first to be different from rusting. Still, as similarity exists in the respect that the white earthy powder,

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formed by the heating, weighs more than the original metal, we are led to believe that in this case also the metal has united with the active part of the air. It has previously been found that not only copper, but also zinc, tin, and lead formed a rust, so that we may state that *most metals*, either by the influence of heat or exposure to moist air, form a *rust*, and that the rusting is accompanied by a gain in weight due to the *abstraction of a portion of the air*, this portion occupying about *one-fifth* of the volume of the air.[1]

The Rusting of Zinc.—In the case of zinc, when heated slightly it becomes coated with a rust, but when heated very strongly it actually burns, forming the *same rust* (p. 31). The zinc therefore may be regarded as forming a link between a metal which rusts, as copper, and one which burns, such as magnesium. This consideration leads further to the idea that there is *no essential difference* between *burning* and *rusting*. Investigations suggested by this idea will be reserved for a later chapter.

34. EXAMINATION OF OTHER METALS.

i. **Preliminary experiments.**—Examine each metal according to the scheme in Chap. II. Also determine the density of each metal.

ii. **Silver.**—(*a*) If possible, examine a sheet of thin silver leaf between two sheets of glass. Observe that when very thin it transmits blue light.

(*b*) Place a silver spoon and an ordinary electro-plated spoon upon a sand bath, as in Fig. 29. Upon the end of each put the end of a wax vesta without any wax with it, or hold the head of a match at the end of each when they are hot. Heat the sand-bath by placing a Bunsen burner under it. The match on the silver spoon will take fire before that on the other spoon.

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[1] The student will do well to observe that this abstraction has only been *proved* (by himself) in a few cases, and the general statement is based here on the analogy of the phenomena. Should anything lead to doubt regarding the general statement, the metal in question should be further investigated.

iii. **Mercury.**—(*a*) Satisfy yourself by trial that, (1) an iron key or lead floats on mercury, (2) mercury does not wet glass, (3) mercury adheres to clean zinc or copper, forming what is called an *amalgam*.

(*b*) Place some mercury in a crucible heated by a small Bunsen flame, and leave it so heated for a few days. At the end of this time note the appearance. Scrape the surface and observe the result.
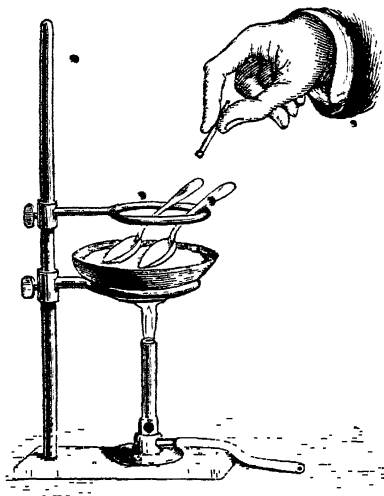

FIG. 29.—Thermal conductivity of silver.

iv. **Gold.**—Examine some thin gold leaf. Place a sheet between two pieces of glass, and observe that when held up to sunlight green light passes through.

v. **Platinum.**—Heat very strongly some platinum wire in a blowpipe flame, and observe that although it becomes white hot it does not fuse.

Silver.—Silver is a white metal about ten and a half times as heavy as water. **It does not tarnish** when exposed to the air, even when heated. It is consequently much used for coinage and for ornamental purposes. It is, however, too

soft to be used by itself, and is generally alloyed with copper. British coins contain about seven and a half per cent. of copper.

Silver conducts heat more readily than any other metal. It is very malleable. When hammered into very thin leaves it is transparent to some constituents of white light. Silver is also very ductile, and can be drawn out into wires of exceeding fineness.

Mercury.—Mercury, or quicksilver, is the only metal which is liquid at ordinary temperatures. Its appearance is familiar to everyone from its frequent use in barometers and thermometers. It is the heaviest liquid known, being $13\frac{1}{2}$ times as heavy as water. When heated in air it slowly tarnishes if the temperature is not too high, whilst if heated sufficiently, it boils, the vapour condensing again to form the liquid. It dissolves many metals, *e.g.* zinc and copper, forming alloys known as *amalgams*.

Not only is it used in barometers and thermometers, but also in the manufacture of looking-glasses; and it is employed in the laboratory, instead of water, for the collection of gases which are soluble in water.

Gold.—Gold is nearly always found native, that is, is found in nature in the pure state. Everybody is familiar with its bright yellow colour and with the circumstance that it is unacted upon by the air. It is more than nineteen times heavier than water. Gold is too soft in the pure state to be used either for coinage or for jewellery, and is always alloyed with copper. This gives rise to the employment of the term *carat*. Pure gold is known as 24- carat gold. The British sovereign, which contains 22 parts of gold in 24 parts of the coin is said to be made of 22-carat gold. Similarly 9-carat gold consists of 9 parts of gold in every 24 parts of the article made of it.

It is the most malleable and most ductile metal known. Gold leaf has been made into sheets so thin that it would

take more than 200,000 of them together to make the thick-
ness of one inch. Gold wire has been made so extremely
thin that two miles of it would only weigh one gram. Thin
gold leaf allows green light to pass through it, so that when
held up to the light it appears green.

Platinum.—Platinum is a rare metal, being about as expen-
sive as gold. It possesses a greyish white colour and a
metallic lustre. Its density is very high, about 21·5; it is
in fact almost the heaviest substance known, being about
twice as heavy as lead. It does not rust nor tarnish, and is
unaltered even when very strongly heated, and owing to this
(and other important properties) it is very largely employed
in chemical laboratories in the form of crucibles. It is very
ductile and may be drawn out into fine wire, this wire being
also commonly used in chemical laboratories.

Use of metals which do not rust.—Although many
metals rust, yet some metals will not do so. These
are of great value for the manufacture of articles which
are required to be permanent and unaltered by the action
of moisture or heat. Some metals, also, as mercury, will
only tarnish or rust with difficulty. A list of all the
metals examined, and of the nature of their rusts, should
be made at this stage. Although rusts of gold, silver,
and platinum are not formed either by the effect of
moisture or of heat, they may be formed by other means,
and it does not follow that the rusts cannot be formed, or
that they do not exist, because the student has been unable
to obtain them in the simple experiments he has performed.

QUESTIONS ON CHAP. IV.

1. Describe experiments which prove that air is composed of at
least two gases

2. How would you show by experiments that only one part of the
air is concerned in (a) burning, (b) rusting?

3. Describe experiments indicating that rusting and burning are
essentially similar operations.

4. When iron is allowed to remain immersed in water it rusts. How could you prove that this rusting is really due to air dissolved in the water?

5. Describe, with the aid of a diagram, an experiment to show that a metal will not rust in inactive air.

6. What conditions are necessary for the rusting of iron, copper, zinc, tin, and lead?

7. What different methods are adopted to preserve iron articles employed out of doors, or used for cooking purposes?

8. Indicate the course of experiments you would carry out, with junior students, in order to determine whether iron in rusting removes the same constituent from air as does copper when being converted into copper scale or calx.

CHAPTER V.

COMBUSTION OR BURNING.

THE experiments performed on the rusting of metals have suggested that burning and rusting are similar processes, at least in the case of metals. It will be well to examine this further, and to investigate closely some particular case of burning. The burning of phosphorus is a convenient case to study, but in all the experiments in which phosphorus is used it must be remembered constantly, that, owing to its extreme inflammability, **phosphorus must never be handled with the bare fingers.** Pieces of phosphorus should be moved from one place to another by a pair of forceps. Whenever it is necessary to cut it, this should be done under water, and the pieces should be carefully dried by blotting-paper before being used.

35. ARE BURNING AND RUSTING SIMILAR PROCESSES ?

i. **Burning of phosphorus.**—(*a*) Read what is said about phosphorus above, and then place a little phosphorus in a test-tube fitted with a good cork. Fix the cork firmly in the test-tube. Hold the test-tube slantingly, by means of a test-tube holder, over a flame for a few seconds, so as to heat the phosphorus and make it burn. When the phosphorus is alight the test-tube should be removed from the flame and the phosphorus allowed to burn. When it will burn no longer, let the tube cool for five or ten minutes. When cool take out the cork under water. Note that the water rises in the tube. Put back the cork and

shake up; test the remaining gas: it extinguishes a burning taper. Measure the volume of the water: it is one-fifth that of the total volume of the closed tube (Fig. 30).

(*b*) Cut a small slit in a large cork, so that the small handle at the top of a crucible lid will fit firmly into it. Place the lid upon the cork, and float the cork upon the water contained in the pneumatic trough. Carefully cut off a piece of phosphorus as large as a good-sized pea, dry it, and place the lump upon the floating lid. The cutting should be done under water. Ignite the phosphorus and slowly place over it a jar or a bottle (Fig. 31). Let the bottom of the bottle either rest on the shelf of the trough or on something suitable, placed on the
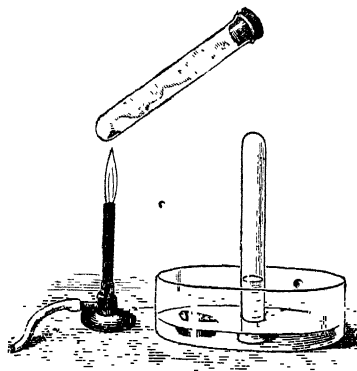
FIG. 30.—Burning of phosphorus.

bottom. Allow the apparatus to remain for a few minutes, when the white fumes will disappear, and the water will be seen to rise up in the jar. Place a greased glass plate
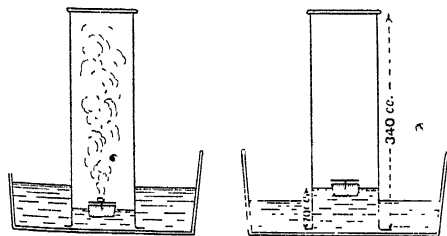
FIG. 31.—Burning of phosphorus in a closed jar over water

under the mouth of the bottle, or jar, and pressing it tightly so that no water escapes. Remove the bottle and set it upright, keeping the glass plate on the mouth. Insert in the jar (1) a lighted taper. (2) a lighted candle. Note that in each case the light is extinguished.

(c) Measure the volume of water in the jar and the total volume it can hold, and determine in this way the percentage of air which has been removed. Compare this with that removed by the iron during rusting, and compare also the properties of the residual air.

(d) Carefully dry a wide glass cylinder and a small crucible. Cut off (under water) a piece of phosphorus about half as big as a pea, and dry it between blotting paper. Using a pair of tongs, place the phosphorus in the crucible, touch it with a hot wire, and quickly put the cylinder over it, as in Fig. 32. A white powder is deposited upon the sides of the cylinder. When the phosphorus has ceased to burn, lift up the cylinder and pour a little water into it. The white powder dissolves with a slight hissing noise.
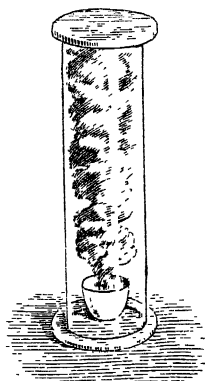
FIG 32.—Collection of fumes from burning phosphorus.
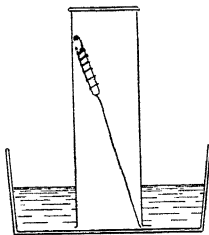
FIG. 33 —Slow combustion of phosphorus.

(e) Place a long stick of phosphorus, fixed to a piece of wire twisted around it, in the bottle used in Expt. 35 (b), and invert the bottle over water (Fig. 33). Allow it to remain for a few days, and again examine. Observe that the water has risen in the bottle. Repeat the experiments, and observe that the phosphorus has slowly taken up one-fifth of the air, and has left the inactive four-fifths. Compare this result with that obtained by the burning of phosphorus (Expt. 35, b) and that obtained by the rusting of iron (Expt. 31).

Different kinds of phosphorus.—·There are two kinds of phosphorus, one called yellow phosphorus (this is the ordinary phosphorus) and the other red, or amorphous, phosphorus. Yellow phosphorus catches fire very easily; the warmth of the hand is quite enough to inflame it. For these reasons it is always kept under water. It is generally bought in the form of sticks, which, when freshly manufactured, are of a yellow, waxy colour. This phosphorus can easily be cut with a knife, but the cutting should always be done under water. However small the piece, it must never be touched with the bare fingers, but always lifted by small tongs or forceps. If this precaution is not taken, the warmth of the fingers may cause the phosphorus to catch fire, and, as it is difficult to shake it off when once alight, the burn which it causes is very severe and dreadfully painful. In all experiments with yellow phosphorus these warnings must be borne in mind. The red form of phosphorus is not so inflammable as the yellow, and has not to be kept under water, but it must be used with care.

Phosphorus readily burns in the air.—It is only necessary to touch a piece of dry phosphorus with a hot wire to make it catch fire and burn. It burns with a dazzling, bright flame, and at the same time dense clouds of white fumes are formed, which spread throughout the room. These facts are noticed until all the phosphorus has disappeared.

Change produced in air by burning phosphorus.—To decide whether phosphorus in burning causes the same change in air as iron does when it rusts, it is best to burn some phosphorus in an enclosed amount of air in a way similar to that which has already been described for an experiment with damp iron. One way to do this is to place a little phosphorus on a cork, or basin, which floats on the surface of water, under a bell jar, or a bottle. After the experiment is over, and the fumes have disappeared,

the water is seen to have risen in the jar, indicating that
there is less gas in the jar than before the phosphorus
was burnt in it.

From what has been previously said, you can understand
at once that phosphorus, in burning, takes out the active
part of the air, and leaves the inactive part behind. So far,
then, the changes which occur when phosphorus burns are
very like those when iron rusts. Some differences will be
studied a little later.

The fraction of the air which disappears as a result of the
burning of the phosphorus in a jar, can be measured
easily enough after the jar has been removed from the water.
As in the case of the rusting of iron, one-fifth of the air is
taken out of it by the phosphorus in burning.

That the gas left behind is really the inactive part of air
can be proved by introducing a lighted taper into the jar or
bottle in which the phosphorus has burnt. The flame is at
once extinguished.

**Phosphorus slowly takes out the active part of the air
without being lighted.**—It has been seen that iron slowly
takes the active part of the air and combines with it to form
rust. And this happens without heating the iron. Will
ordinary phosphorus do the same when it is not alight?
This question, too, is easily answered by a simple experi-
ment. When a piece of clean phosphorus is enclosed
in a quantity of air over water, the rapid changes just
described take place slowly. The only difference in the two
cases is the rate at which the active part of the air is taken
out. Burning phosphorus combines with the active part
very quickly ; if not burning, only slowly. Still, given time
enough, ordinary phosphorus will remove all the active part
of air without burning, and at the end of the experiment it
will be found that again one-fifth of the air has disappeared.

**Properties of the substance formed when phosphorus
combines with the active part of the air.**—We have, as

yet, only noticed that the substance which is formed when phosphorus removes the active part of the air is a white snow-like powder, which corresponds with the rust formed when iron is exposed to air. With a little care, the quick disappearance of the white material which forms the fumes of burning phosphorus can be prevented. All that need be done is to burn a piece of dry phosphorus in a dry vessel. In these circumstances, the white fumes settle down on the inside of the vessel in the form of a snow-like solid. But the white powder has so great an attraction for water that as soon as the vessel is opened it extracts the moisture from the air, and, first becoming moist, is quickly replaced by drops of liquid. If water is put into a test-tube or other vessel in which dry phosphorus has been burned, the white powder rapidly dissolves with a hissing noise, like that noticed when water comes in contact with hot iron.

Results of the experiments. — All these experiments completely verify the idea suggested by the burning of magnesium, viz., that burning and the rusting of metals are essentially similar processes. The conclusions derived from these experiments are, that the phosphorus takes up, or unites with, a part—the active part—of the air, forming a white powder, and leaves the inactive air. If this is so, then it is clear that the white powder should weigh more than the phosphorus from which it is formed, and further, metals should not be able to rust in the residual air left after the burning of phosphorus. (See Expt. 32.) It will be well, therefore, to perform these two tests.

36. BURNING AND RUSTING ARE SIMILAR PROCESSES.

i. **Change of weight of phosphorus during combustion.** ---Obtain a hard glass tube, BA in Fig. 34, pack the drawn-out end with asbestos fibre ; weigh carefully. Place in it a small piece of dry phosphorus (about 0·2 gram) and again weigh. Connect up as

shown in Fig. 34, the end A being attached to a suction pump[1] which can be fitted to a tap. This draws a current of air through the tube, and, in order to dry the air, it is first made to pass through the test tube C, containing strong sulphuric acid. Warm the phosphorus slightly—it ignites. Take away the burner and allow it to burn, the fumes being stopped by the
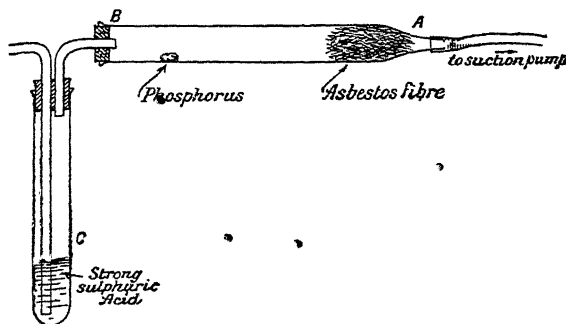

FIG. 34.—Change of weight during combustion of phosphorus.

asbestos. A quantity of red deposit is also found (another form of phosphorus), and this may be got rid of by stiongly heating. When cool, disconnect, and again weigh the tube. It is found, if care has been taken, that the phosphorus gains in weight.

ii. **Metals will not rust in residual air.**—Burn some phosphorus in a *large* bottle or jar in the manner explained in Expt. 35 (b). When combustion is complete remove the jar as before and repeat Expt. 32, ii., but using this residual air instead of the air after the rusting of iron. You will find that the copper will not rust in this residual air.

Substances gain in weight during combustion. The residual air is the same as that left ꞌafter rusting.—It is seen that these experiments completely confirm the con-

[1] If suction pumps are not available, the current of air may be obtained by filling a *large* bottle with water, and closing it by a cork or india-rubber bung, through which pass two tubes : one, the short one, only passing just through the cork ; the other long, one passing to the bottom of the bottle. The outer end of the long tube is fastened to a long piece of india-rubber tube dipping into a sink, etc., *below* the bottle. When the water is sucked into this india-rubber tube it continues to syphon off, and draws a current of air through tubes, etc., fixed to the short tube (Fig. 35).

clusions derived from the burning of phosphorus. When phosphorus burns in a tube, so arranged that the white fumes formed do not escape into the air, but remain in the tube, it is found that the weight of the tube, etc., is greater after the burning than it was before, so that the white powder formed weighs more than the phosphorus itself. The residual air left after phosphorus burns in a closed space, will neither support combustion nor allow metals to rust in it ; it occupies four-fifths of the volume of the air, so that we conclude it is the *same inactive air* as that left when metals rust.

Screw Clip

Air

A

B

Water
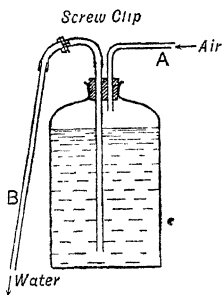
FIG. 35.—A form of aspirator.

These combustion experiments, therefore, again indicate that air consists of *two parts*, one the *active part*, occupying *one-fifth* of the volume of the air ; the second an *inactive part*, occupying four-fifths.

Further, we may state that *rusting* and *burning* both consist in the *union* of the *metal* or *burning body* with the *active part* of the *air*.

QUESTIONS ON CHAP. V.

1. Describe experiments which prove that during rusting and burning an increase of weight occurs.

2. How would you show that the slow combustion of phosphorus is a process similar to its rapid burning?

3. Write a short account of the different methods you have seen used in order to make a dull fire burn brightly, and show how in each case the result is achieved.

4. How would you prove that when phosphorus burns, the product formed weighs more than the original phosphorus?

5. What reason have you for considering that the gas left after the burning of phosphorus is the same as that left after the rusting of iron?

CHAPTER VI.

CONSTITUENTS OF THE AIR: OXYGEN AND NITROGEN.

The active constituent of air.—We must now endeavour to obtain and examine the active part of the air which disappeared during the rusting of iron, etc. Knowing that it is now present in the rust, the most evident plan would be to endeavour to obtain it from this source. The ease with which iron rusts, that is, the readiness with which it takes up the active part of the air, should indicate that it would probably be very difficult to obtain it from this source, and that some other rust which is more difficult to prepare would probably be better for our purpose. The most convenient is the rust of mercury, which is a red powder not easily formed (Expt. 34, iii.).

37. SEARCH FOR THE ACTIVE PART OF AIR.

i. Heating of mercury rust.—Place some of the mercury rust (known as red oxide of mercury) in a tube of hard glass closed at one end, and heat strongly. Notice the darkening of the powder, also the dark deposit which collects round the inside of the tube above the powder. Place into the tube a splinter of wood which has just been extinguished, and is still glowing. Note that it glows more brightly, or even bursts into flame. Allow the tube to cool, and notice that the powder returns to its original colour. With a piece of wood, or glass, scrape off the dark deposit from the walls of the tube. It is seen to be bright, metallic quicksilver or mercury.

ii. **Changes produced by heating lead in air.**—Heat a few pieces of clean lead in an open crucible (Fig. 36). When the lead has melted, stir the liquid metal with a stout iron wire. Notice the formation of a powdery scum upon the lead. Observe that the colour of the powder is darker when hot. Let the crucible cool. Notice that it now contains a yellow powder in addition to the unchanged lead. By strongly heating this powder its colour changes again, and it becomes *red lead*.

FIG. 36.—Heating of lead.

iii. **Gas produced by heating red lead.** —Place a little red lead in a hard glass tube, and strongly heat the tube as in Fig. 37. Notice that the red lead undergoes a change of colour. Into the tube insert a glowing splinter. Observe that the splinter is rekindled. Why is this?
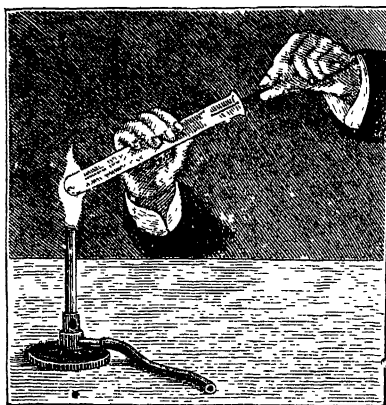
FIG. 37.—Heating of red lead.

Red oxide of mercury—the action of heat on it.—Red oxide of mercury, or, as it is sometimes called, red precipitate, can be obtained by heating metallic mercury for a considerable time in the air, when the oxide forms as a red scum upon the surface of the metal.

When red oxide of mercury is strongly heated, it darkens

in colour, and a dark deposit of mercury collects above the powder round the inside of the tube in which the heating is done. When a glowing splinter of wood is plunged into the tube it bursts into flame.

What does this experiment teach? It shows that by heating the rust of mercury we obtain mercury itself, and also a gas in which wood burns more brightly than in air. The gas thus obtained was taken out of the air when the mercury rusted, and it is given up when the rust is heated. As the gas which causes rust to form is the active part of air, it supports combustion very vigorously, for we have found that the other four-fifths of the air is inactive, not supporting combustion at all, and hence acting as a diluent.

It may further be proved, by weighing, that the weight of the original mercury is equal to that left after the experiment, provided that all the rust is decomposed and no mercury is lost. It may also be proved that the volume of the gas abstracted from the air, during the rusting of the mercury, is exactly equal to the volume of gas evolved, when the mercury rust is heat. This shows that the gas escaping from the rust is the same gas as that taken from the air.

The compounds which lead forms with the active part of the air.—When lead is heated in contact with the air, a yellow powder, which is much darker in colour when hot, is formed. If the heating is continued long enough, all the metal is changed into powder. The change takes place fairly easily, so that, from previous reasoning, you would conclude that it is probably difficult to get the active part of the air again from this powder. And this is so. But it is found that, when some of the yellow powder is heated for a long time at the temperature at which lead melts, it slowly takes up still more of the active part of the air, and changes in colour, becoming red. The first powder obtained, which is yellow, is in some states called *litharge*; the second, red

powder is known as *red lead*. It is easy to get the second portion of the active part of the air again from the red lead.

How the active part of air is obtained from red lead.—When red lead is heated, it changes in colour, and if the heat has not been great, it regains its original red colour when allowed to cool. But if strongly heated, the red lead gives up some of the active part of the air which it contains, and is reconverted into litharge. The amount of the active part of the air which it thus gives up, on being heated, is the second quantity referred to in the last paragraph, which is taken up slowly when the heating of lead is continued for a long time. If red lead is strongly heated in a tube, and a glowing splinter of wood is pushed down the tube, the splinter bursts into flame and continues to burn brightly. The active part of air has been obtained alone, and supports burning very strongly.

The active part of the air is called oxygen.—As it will be more convenient in the future to speak of the active part of the air by the name chemists use for it, we may state here that it is always called *oxygen*. We may hence speak of the change which takes place when the oxide of mercury is heated thus :—*Oxide of mercury* on heating forms *oxygen* and *mercury*.

38. PREPARATION AND PROPERTIES OF OXYGEN.

i. **Oxygen from potassium chlorate.**—Place a little potassium chlorate or chlorate of potash (which is the same thing) in a test-tube, and heat it. Observe that the powder crackles, melts, and gives off a gas. Test by a glowing splinter of wood, and see that the gas behaves like oxygen, the active part of the air.

ii. **Preparation of a small quantity of oxygen.**—Powder some crystals of potassium chlorate, and mix the powder with a little manganese dioxide (sometimes called pyrolusite). Heat some of the mixture in a test-tube, as in the last experiment. Observe by putting in a glowing splinter that oxygen is given off. Notice that in this case there is no melting, and the gas comes off more readily.

iii. **Preparation and collection of oxygen.**—Into a hard glass tube, closed at one end, fit an india-rubber stopper, with one hole in it, through which a tube, bent as in Fig. 38, is passed.
The other end of this tube, called the delivery tube, dips under water in a trough. Mix together some potassium chlorate and manganese dioxide, as in the previous experiment, and place the mixture in the tube. Support the tube and delivery tube as shown in Fig. 38. Fill several bottles with water, and invert them in the trough. Gently warm the tube, and place one of the bottles of water over the end of the delivery tube.
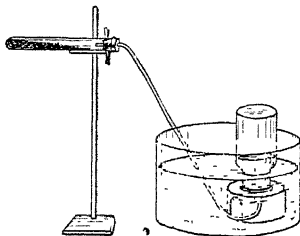
Fig. 38.—Preparation and collection of oxygen

As the oxygen is driven off, it displaces the water and gradually fills the bottle. When the bottle is full of oxygen, cover its mouth with a greased glass plate, and lift it out of the trough. In this way fill five or six bottles with oxygen.
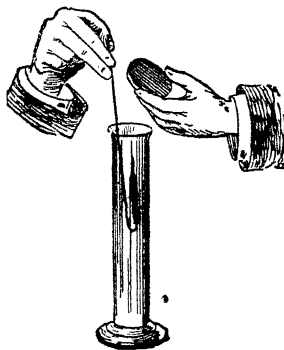
FIG. 39.   FIG. 40.—Burning of a candle in oxygen.

Caution.—*Be careful not to take away the burner from under the hard glass tube before removing the delivery tube from the trough.*

iv. **Physical character of oxygen.**—Take one of the bottles, and examine the gas inside as far as you can by looking at it. Then lift off the glass plate and smell the gas. Draw a little into the mouth, and notice that it has no taste.

v. **The burning of a candle in oxygen.**—Attach a piece of stout wire to a wax taper, as shown in Fig. 39, and having lighted the taper plunge it into another of the jars of oxygen (Fig. 40). Notice that it is *not extinguished*, but continues to burn, but with a *larger* and *brighter* flame.

vi. **The burning of charcoal in oxygen.**— Into another jar of oxygen thrust a splint of wood, red-hot at the end, or a piece of red-hot charcoal placed in a deflagrating spoon (a small upturned iron spoon with a long handle) (Fig. 41). Note the brilliancy of the combustion. Test the gas left in the bottle with (1) some moist, red litmus paper, (2) some moist, blue litmus paper. Note it changes the colour of the latter to a reddish purple.

vii. **The burning of phosphorus in oxygen.**— In another jar burn a piece of phosphorus, about half the size of a pea, contained in the deflagrating spoon. Note the great brilliancy of the combustion and the dense white fumes. Add water and shake—the fumes dissolve. Into the solution put a blue litmus paper. Observe that it is turned *red*

viii. **The burning of sulphur in oxygen.**— Perform, with another jar, a similar experiment with sulphur. There are few fumes, and a strongly smelling gas is obtained, also soluble, turning blue litmus red.

FIG. 41 —Deflagrating spoon in jar.

ix. **The burning of magnesium ribbon in oxygen.**—Ignite a small piece of magnesium ribbon and hold it by means of crucible tongs in a jar of oxygen. Notice the white solid formed. Test its solubility in water, and show that unlike the previous products, it will not turn blue litmus solution red, but will turn red litmus solution blue.

x. **The burning of sodium in oxygen.**—Put a small piece of sodium [1] in a *dry* deflagrating spoon, light the sodium by heating it strongly, and put it alight into another jar of oxygen.

[1] Great care must be taken when using sodium, which must never be allowed to touch damp materials. It is kept under naphtha until used, and should never be handled with the fingers. It should be dried by blotting paper when taken from the bottle and cut with a clean knife, the pieces not used being immediately replaced in the bottle.

Observe the fumes formed.   Dissolve these in water, and try the effect of the solution on litmus paper.   It does not turn the blue colour to red, but has exactly the reverse effect, and changes red litmus to blue   Feel the water ; it has a soapy feel.

xi. **The burning of iron in oxygen.**—Obtain a piece of iron wire (a thin steel watch spring will do), and dip one end into a little melted sulphur, and when the sulphur is burning place the wire in another jar of oxygen.   Observe that the sulphur burns and also starts the combustion of the iron, which continues to burn with a brilliant shower of sparks.   After the burning has ended, observe that a quantity of an insoluble solid (iron rust) has been formed.

xii. **Substances when cold do not burn in oxygen.**—Place any of the above materials *cold* into a jar of oxygen and note that there is no effect.   They do not burn.

Other substances besides rusts of metals give up oxygen when heated.—It has been found by experiment that a number of substances give off oxygen when heated.   The one most commonly used is a white, crystalline solid, called potassium chlorate.   If a crystal of this substance is placed in a test-tube and heated in the flame of a laboratory burner, it first crackles, then melts, and by and by begins to give off bubbles of oxygen gas.   The gas given off appears to be oxygen, because when a glowing splinter of wood is inserted into the tube it is immediately re-kindled, bursting into flame and burning brightly.

Oxygen is given off more easily if the potassium chlorate is mixed with manganese dioxide.—To obtain all the oxygen from potassium chlorate requires a considerable amount of heat, when the substance is heated alone   If, however, the chlorate of potash is first mixed with certain other substances, such as manganese dioxide, it is found possible to obtain the oxygen with very much less trouble. The gas is given off at a much lower temperature, and more readily in every way.   It is the custom, therefore, when it is desired to prepare considerable quantities of oxygen in the laboratory, to use a mixture of potassium chlorate and manganese dioxide, which is often known as *oxygen mixture.*

Preparation and collection of oxygen.—The apparatus shown in Fig. 38 is very convenient for preparing oxygen from oxygen mixture. A tube closed at one end is fitted with an india-rubber stopper, or a good sound cork, through which a hole has been bored. A small glass tube, conveniently bent, fits tightly into the hole, and is so arranged that its other end dips under water in a trough. Oxygen mixture is placed in the closed tube. When this mixture is heated, a gas is given off. To collect this gas, several bottles should be filled with water and inverted in the trough. When these bottles are ready, the closed tube is gradually warmed with a laboratory burner, and after a little while one of the bottles is placed over the open end of the tube in the trough. The gas which passes down this delivery tube, being lighter than the water, of course rises up through it into the bottle. In this way the water is gradually pushed out of the bottle, and its place taken by the gas.

This plan of collecting gases which do not dissolve, or dissolve only to a small extent, in water, was devised by Priestley, the chemist who first obtained oxygen. It is called *collecting over water*. When the bottle is full of the gas, and while its mouth is still under water, it is covered with a greased glass plate, which is held tightly to the mouth by the left hand, and the bottle lifted out and placed on the table with the right. The gas is now ready for testing.

Properties of oxygen.—If one of the bottles of gas, collected as described, is allowed to stand for a minute or two and then examined, several of the characters of oxygen can be made out. To begin with, oxygen is an invisible gas. It has neither smell nor taste. The method by which it was collected is also a sufficient proof of the fact that it is either (1) insoluble in water, or is (2) soluble to only a very slight extent. The experiments previously performed on rusting (p. 59) show that water ordinarily contains a small

quantity of oxygen, so that the gas is slightly, but only slightly, soluble in water.

Oxygen has no effect on substances like sulphur and carbon when they are at the same temperature as the room, but if these elements are heated to the point of ignition the oxygen *combines* with them very readily, causing them to burn vigorously.

If, before putting a piece of charcoal into oxygen, it is made red-hot in the flame of a laboratory burner, or a spirit-lamp, it no sooner comes into contact with the oxygen than a change is noticed. The piece of charcoal begins to burn very brilliantly, and bright sparks are given off. As the burning proceeds, the piece of charcoal gets smaller and smaller.

Similarly with sulphur; even when a piece is held in a flame, it only catches fire with difficulty, and burns with a very pale, blue flame, giving but little light. But if, while it is still alight, it is transferred to a jar of oxygen, a difference is at once observed. The flame gets larger and brighter, and is of a beautiful lavender colour. It remains until all the sulphur, or all the oxygen, or both, have been used up.

The flame with which phosphorus burns in oxygen is dazzling in its brightness. In fact, the burning is so intense that it is painful to some eyes to watch it. This is particularly the case when the experiment is performed in a dark room.

Magnesium also burns with great brilliancy in oxygen, and forms thick clouds of a white powder. Sodium does not burn very readily, owing to the fact that when heated, or even when merely exposed to the air, it becomes coated with a scum which cannot burn. It may, however, be made to burn, and if immediately plunged into a jar of oxygen it continues to burn very brightly. Even iron wire will burn in oxygen if heated to redness, which may be done by a small tip of burning sulphur. On account of these pro-

perties oxygen is spoken of as a very vigorous *supporter of combustion*.

Oxides and their properties.—In all these cases of burning, both the substance burnt and the oxygen disappear and a new product is found. They have in fact combined to form a compound, which we term an *oxide*, the name given to a compound of oxygen with some other element. These oxides differ in their appearance and in their properties, as is readily seen if the contents of the bottles are examined after the burning.

Carbon.—Thus, the bottle in which the carbon was burnt contains a colourless gas, which does not support combustion, has no marked smell, and turns moist, blue litmus paper to a reddish purple colour. This gas is an *oxide of carbon.*

Sulphur.—The sulphur when it burns forms a gas with a very strong, pungent odour—*oxide of sulphur*—which dissolves in water, and turns blue litmus paper to a bright red. (It was noticed that white fumes are produced when sulphur burns in oxygen. These, however, are partly unchanged sulphur and partly another product which is formed in smaller quantities, and may be reserved for study in some later course)

Phosphorus.— Phosphorus forms a white powder—*oxide of phosphorus*—which is very soluble in water, the solution behaving to blue litmus in the same way as did the solution of oxide of sulphur, *i.e.* it changes the colour of the litmus to a bright red.

Magnesium.—This metal also forms, by combustion, a white powder—*oxide of magnesium*—which differs from the oxide of phosphorus in being almost insoluble in water. That it is slightly soluble, however, is seen by the fact that when shaken up with water, the liquid then has an action on litmus paper. This action, however, is different from that of the previous oxides—instead of turning blue litmus red, it has the reverse action and turns red litmus blue.

Sodium.—The white *oxide of sodium* is very soluble in water, and the solution has a peculiar soapy feel, whilst it has a very powerful action on litmus paper ; an action of the same kind as that of the oxide of magnesium, that is, it turns red litmus blue.

Iron.—The *oxide of iron* is a dark, reddish brown powder, which is insoluble in water.

Different classes of oxides.—It is seen that the oxides formed may be divided into certain classes according to their action on litmus paper. Some have the power of changing blue litmus red—others turn red litmus blue. The first of these are termed **acids**, and if the various compounds in the laboratory labelled, *acid* be examined, it will be seen they all have this effect on litmus paper. The second class of oxides act on litmus in the reverse way, and these are termed **alkalis**. Some oxides have no action on litmus paper and are termed *neutral.*

We may now write

Phosphorus and oxygen form oxide of phosphorus (**acid**),
    sodium and oxygen form oxide of sodium (**alkali**),

and so on

It was owing to the formation of acids by the solution of some oxides that the gas received the name *oxygen*, which means acid-producer. It must be remembered, however, that *all* oxides are not acids, some being alkalis.

39. IDENTITY OF GAS FROM POTASSIUM CHLORATE AND MAGANESE DIOXIDE WITH THE ACTIVE PART OF THE AIR.

i. **Burning in air.**—Burn sulphur, phosphorus, magnesium, sodium, and charcoal in bottles filled with ordinary air. Examine the products exactly as the products of their combustion in oxygen were examined. Note that the properties are identical.

Identity of active gases examined—The objection might be raised that although the gas obtained from potassium

chlorate and manganese dioxide is an active gas, yet it may not be the *same* active gas as that present in the air, which is also obtained from the oxide of mercury, for it is quite possible that there are more than one active gas.   This objection is well founded, and the identity of this gas with that in the air should be proved.   To do this, we may endeavour to ascertain if the products of combustion in the previous sets of experiments are identical with the products of combustion in air, when, as we have seen, the burning substance only combines with the active constituent of the air—the oxygen. We find that the products are really identical, sulphur forming the pungent smelling, acid gas, sodium the white, alkaline powder, etc.

We may conclude, therefore, that the gas formed by heating potassium chlorate, or its mixture with manganese dioxide (called oxygen mixture), is identical with the active part of the air which we call oxygen.

The inactive part of the air has been previously obtained, but, it will be well to revise and repeat the work, comparing the properties of the inactive part with those of the active constituent—oxygen.

40. THE INACTIVE PART OF AIR.

i. **The inactive part of air.**—Repeat the experiments of allowing iron to rust in an enclosed quantity of air over water, and of burning phosphorus under a jar of air over water, and satisfy yourself that the gas left behind (*a*) extinguishes a flame; (*b*) has no action on a litmus paper.

ii. **Air can be obtained again by mixing oxygen with the inactive part of air.**—Again allow iron to rust in an enclosed amount of air.   When the volume of gas in the bottle has ceased to diminish, remove the muslin bag containing the iron.   Place the delivery tube from an oxygen apparatus, similar to that used in the Expt. 38, iii., and by heating the hard glass tube bubble oxygen into the bottle until it is again full of gas.   Cover the mouth of the bottle with a piece of cardboard, lift out the bottle, and test the gas it contains with a burning taper; it behaves just like air.

The inactive part of air.—The gas which is left in a bottle of air after iron has rusted in it, or phosphorus has burnt in it, or in which any one of the instances of burning which have now been studied has occurred, will no longer allow a candle or taper to burn in it. This is one reason why it is called the inactive part of air. The inactive part of air does not affect damp iron at all ; that is, the iron does not rust when put into it. The name by which this gas is known to chemists is *nitrogen*, and the effect of its presence in the air is that the activity of the oxygen is thereby diminished, toned down, or weakened. Nitrogen not only will not allow things to burn in it, but it will not itself burn, or, as is more commonly said, it is *incombustible.* There are other substances in the air besides the active and inactive constituents which have now been described, but in this chapter, since their amounts are small, their study would be out of place.

Properties of nitrogen.—This gas nitrogen is an example of a very inactive, or inert, substance. It can only be made to combine with other substances with great difficulty—in fact, with most other things it cannot be made to combine at all. Not only is nitrogen incombustible, and a *non-supporter of combustion*, but it also does not support life, and a mouse put into the gas dies very quickly—this is because of the absence of oxygen.

If the number of negative properties of nitrogen is borne in mind, and if we think of this side by side with the very active qualities of oxygen, it is not at all difficult to understand that its presence in the air serves the purpose of weakening the oxygen, and thus making combustion less intense.

QUESTIONS ON CHAP. VI.

1. Describe briefly the sequence of experiments which indicate the existence of oxygen in air, and give a mode of preparation of this gas.

2. How is oxygen most conveniently prepaied? Give an account of its characteristic properties.

3. What are oxides, and how may they be obtained?

4. How can nitrogen be obtained from the air, and what are its chief properties?

5. How would you find the quantity of oxygen and nitrogen present in a given volume of air?

6. By the combustion of different substances various oxides may be produced. Point out how by the action of water these oxides may be divided into two classes, and give examples.

7. How would you find whether a bottle of oxygen contained any admixed nitrogen?

8. Describe the appearance and properties of the oxides produced by the burning of magnesium, sodium, sulphur, phosphorus, and carbon

9. Describe two soluble oxides and three insoluble oxides.

CHAPTER VII.

QUANTITATIVE STUDY OF THE ACTION OF HEAT ON POTASSIUM CHLORATE.

THE heating of potassium chlorate and manganese dioxide should also be done quantitatively, as the numbers obtained will be of use in later work.

41. QUANTITATIVE CHARACTER OF THE ACTION.

i. **Volume and weight of oxygen from 100 gms. of potassium chlorate.**—Procure a test-tube and put in it a little manganese dioxide (previously *well dried* by heating), and place near the mouth a plug of glass wool, or asbestos ; weigh. Add potassium chlorate and put back the glass wool ; weigh again. Connect up as in Fig. 42 and heat until no more oxygen is evolved ; measure the water driven out of A into the graduated jar B. This is equal to the volume of gas evolved from the potassium chlorate and manganese dioxide. Allow the tube to cool and weigh again.

Enter your results thus :

Wt. of tube, manganese dioxide, and potassium chlorate,	36·85 gms.
Wt. of tube and manganese dioxide,	35·32 gms.
1. Wt. of potassium chlorate,	1·53 gms.

Wt. of tube, manganese dioxide, and potassium chlorate,	36·85 gms.
Wt. of tube, manganese dioxide and potassium chlorate, after heating,	36·25 gms.
2. Wt. of oxygen evolved,	0·60 gm.
3. Volume of oxygen,	430 c.c.

We thus find, if the experiment is carefully performed, that from 100 gms. of potassium chlorate we should obtain $\dfrac{100}{1\cdot53} \times \dfrac{0\cdot60}{1}$ gms., $i.e$ 39 gms. of oxygen, and further, that 430 c.c. of oxygen weigh 0·60 gms., that is, 1 litre weighs 1·39 gms. We
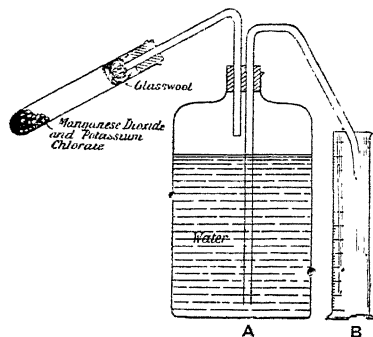
FIG. 42.—To illustrate Experiment 41.

have not here, however, corrected the volume of gas for the pressure and temperature, and this should be done, the barometric pressure and an temperature being read at the time of the experiment, and the volume of gas corrected for the pressure 760 mm. and temperature 0° C. The weight of 1 litre of oxygen will be then found to be slightly over 1·4 gms.[1] The absolute density is hence also known, and is 0·0014 gms. per c.c

In the above experiment it will be noted that the weight of manganese dioxide was not taken, so that in different experiments the quantity has probably varied within wide limits. It will be found nevertheless that the weight of oxygen varies only slightly and within the range of experimental errors.

Oxygen from oxygen mixture—Quantitative results.— When oxygen is prepared from "oxygen mixture," it is found that if the mixture is heated so that the evolution is complete, the weight of oxygen evolved is about 39 grams for every 100 grams of potassium chlorate employed, and that this quantity is independent of the quantity of manganese dioxide added. This experiment

[1] If found more convenient to do so, the experiment may be divided into two, $i.e.$ two separate experiments may be performed, (1) to obtain the weight of oxygen, (2) to obtain the volume of oxygen evolved from 100 gms. of potassium chlorate. In this case, too, the weight of a litre of oxygen should be found.

appears to indicate, therefore, that although the manganese dioxide assists the evolution of the oxygen, as has been already observed (Expt. 38, ii.), yet it does not itself furnish any of the oxygen, as might perhaps have been suspected.   For, if the manganese dioxide evolved oxygen as well as the potassium chlorate, a higher percentage of oxygen —calculated only on the weight of potassium chlorate— would be obtained when the quantity of manganese dioxide was great than when it was small.

By also finding the volume of oxygen evolved per 100 gms. of potassium chlorate, the weight of a litre of oxygen may be obtained.   In this way, it is found that a litre of oxygen weighs about 1·4 gms.   Inasmuch as one litre of water weighs 1000 gms., oxygen is only $\frac{1\cdot4}{1000}$ of the weight of water, that is about one-seven-hundredth.   This weight is, however, by no means negligible, and if each litre of oxygen weighs nearly one and a half grams it is evident that the oxygen in even a small sized room must weigh a considerable amount, and that the weight of oxygen in the atmosphere must be enormous.

42. EXAMINATION OF THE RESIDUE LEFT AFTER HEATING OXYGEN MIXTURE.

We must now return to the tube which contains the residue of the heated mass in Expt. 41.

i. **Examination of residue.**—Heat well, till no more oxygen is given off.   Boil up the mass with water and filter it.   The manganese dioxide remains apparently unchanged on the filter paper. Evaporate the solution to dryness, and there remains a white solid which is not potassium chlorate, and tastes something like salt.

ii. **Weight of residue.**—Weigh a test-tube with some manganese dioxide and glass wool or asbestos as in Expt. 41.   Add some potassium chlorate, mix, and again weigh.   Heat until the whole of the oxygen has been evolved.   Allow the tube to cool, and then boil the contents three or four times with water, filtering each time and allowing the filtrate to pass into an

evaporating dish previously weighed. When all has been filtered, again pour through the filter paper a little hot water. Note the apparently unchanged manganese dioxide on the filter paper. Evaporate the contents of the evaporating basin to dryness on a water bath, and again weigh, observing the white potassium chloride in it. Calculate the weight of potassium chloride left from 100 gms. of potassium chlorate and observe that this together with the weight of oxygen, previously found to be evolved from the same quantity of the compound, makes up (within experimental errors) 100 gms. Consider this, and draw your own conclusions.

The results may be entered in the following manner :

Weight of tube with manganese dioxide and
    potassium chlorate, - - - - - 35·946 gms.
Weight of tube with manganese dioxide, -   -   34·235 gms.

    Weight of potassium chlorate, -   -   -   1·711 gms.

Weight of evaporating basin and potassium
    chloride, - - - - - - 55·859 gms.
Weight of evaporating basin, - - - - 54·815 gms.

    Weight of potassium chloride, - - - 1·044 gms

From 1·711 gms. potassium chlorate 1·044 gms. potassium chloride remain.

From 100 gms. potassium chlorate $\dfrac{1·044 \times 100}{1·711}$ gms. potassium

chloride remain.   = 61 gms.

A residue of potassium chloride and unchanged manganese dioxide remains after heating oxygen mixture.— The residue left in a tube in which oxygen mixture has been heated consists of a hard black mass. If boiled with water, it is found that a large quantity of this mass remains undissolved, in the form of a black powder, which, when the liquid is filtered, remains on the filter paper. It appears to be the manganese dioxide, which is apparently unchanged in the action.

The liquid on evaporation, however, leaves a white solid which is not potassium chlorate, because no oxygen is

evolved when it is heated. It possesses a taste somewhat resembling that of salt, and it is called *potassium chloride.*

It would be unwise, however, to conclude that the manganese dioxide is unchanged without in some further way testing to see if this is really the case. The Experiments 41 and 42, i., both indicate that it is so, and further experiment confirms these results. For, if the residue left after heating a known weight of potassium chlorate be boiled with water and filtered, and the filtrate evaporated to dryness in a weighed evaporating basin, it is found that the weight of potassium chloride left, added to the weight of oxygen evolved, gives a result exactly equal to the weight of the potassium chlorate. The oxygen has therefore been evolved from the latter compound solely, and we may conclude that the black powder on the filter paper is really unchanged manganese dioxide. We may therefore write: *Potassium chlorate when heated, decomposes into potassium chloride and oxygen.*

The manganese dioxide did not change ; what, then, was its rôle ? This question is one which was for a long time explained by calling the manganese dioxide a *catalytic agent,* and such actions, viz., those accelerated by the presence of a substance, which itself does not change, were called *catalytic actions.* It is most probable, however, that the manganese dioxide passes through a series of changes, but the final one leaves it in its original condition.

QUESTIONS ON CHAP. VII.

1. A sample of potassium chlorate contains potassium chloride. How could you estimate the quantity of each in it ?

2. What reason have you for believing that in the preparation of oxygen from ordinary oxygen mixture, the gas is only obtained from the potassium chlorate ?

3. A mixture of potassium chlorate and manganese dioxide is given to you. How could you estimate the quantity of potassium chlorate in it ?

4. On heating chlorate of potash it was found that ·535 gramme of the salt gave 160 c.c. of oxygen, measured 15° C. and 730 m.m. pressure; the residue after heating was found to weigh ·33 gramme ; from these data calculate the weight of a litre of oxygen at N.T.P.

CHAPTER VIII.

ACTION OF ACIDS ON METALS. HYDROGEN.

It has been found that certain oxides dissolve in water to form solutions, which have been termed **acids**. These solutions are characterised by the power they possess of turning blue litmus red, as well as by other properties, which the student will realise at a later stage of his study of chemistry. Some of these acids are of great importance to the chemist, and are in frequent use in the chemical laboratory. The most important are probably those termed *sulphuric acid*, or *oil of vitriol*; *nitric acid*, or *aqua fortis*; and *hydrochloric acid*, also called *muriatic acid*, and *spirits of salt*. These compounds will, in the following work, be always spoken of by the first of the above names, which, in each case, is that by which the acid is always known to chemists. They are all *poisonous* and all attack and destroy clothing, etc., so that great care is required in using them.

43. PROPERTIES OF ACIDS.

i. **Sulphuric acid.**—Take a very small quantity of sulphuric acid in a test-tube. Smell it. Add some water—about an equal quantity—and notice that the tube becomes hot. Place some blue litmus paper in the solution. Pour a couple of drops of the strong acid in a crucible, heat, and observe the thick white fumes formed. Put a drop on to some filter paper or wood, and notice that the paper. etc., is charred. Compare the properties of this ·liquid with those of the liquid you found to be formed when green vitriol was heated.

ii. **Nitric acid.**--Take a small quantity of nitric acid in a test-tube, and, as in the case of sulphuric acid, smell it, and observe the action of water and of heat upon it. In the latter case, observe carefully the coloured fumes evolved by heating, and that the liquid itself becomes coloured. Place a little of the acid on some cloth, or a piece of wood, and notice the yellow stain produced. Observe the action of dilute acid on litmus paper.

iii. **Hydrochloric acid.**--Examine similarly hydrochloric acid. Observe its sharp smell, which is more marked when the liquid is heated. Observe the action of the dilute acid, or of the fumes, on litmus paper.

Sulphuric acid is an odourless, thick, oily liquid, the density of which is 1·85 grams per c.c. When mixed with water, considerable heat is produced and the mixture becomes very hot. Sulphuric acid, indeed, absorbs water very readily from moist substances, and is on this account often used for drying moist gases. Thus, the oxygen prepared in Expt. 38, iii., was collected over water, and would, therefore, certainly contain a quantity of water vapour. If the gas, however, is bubbled through sulphuric acid the water vapour is removed, and the *dry* gas obtained. This method of drying gases will be employed frequently during later work.

Sulphuric acid is also characterised by its power of charring vegetable matter, and by the thick, suffocating fumes it emits when strongly heated. If pure, it is colourless, but if any organic matter gets into the acid and chars, the acid becomes dark in colour. On this account strong sulphuric acid frequently has a slight brown colour.

Nitric acid is a colourless liquid if pure. The ordinary acid is, however, frequently of a yellow colour. It has a choking smell, and when heated boils, and gives off choking fumes of a reddish colour, becoming at the same time darker in colour itself. It stains organic matter yellow and has a very powerful similar action on the skin.

Hydrochloric acid is also a colourless liquid with a pungent smell which cannot be confounded with that of

nitric acid.   Its properties are not as striking as those of the other two acids, but. like them, it is a strong acid and *very great care* must be taken in its use.   All these acids, even when very dilute, give the characteristic reaction with blue litmus paper.

44. ACTION OF ACIDS ON METALS.

i. **Action of acids on iron.**—Place some iron filings in each of three test-tubes and add to (1) dilute sulphuric acid, (2) dilute nitric acid, (3) dilute hydrochloric acid.   Observe carefully the action, if any, in each case.   If any gas appears to be evolved, as evidenced by effervescence in the liquid, test the properties of this gas with regard to its action on (*a*) a glowing splinter of wood, (*b*) a lighted match.

ii. **Action of acids on other metals.**—Repeat the last experiment. but using (1) zinc, (2) copper, (3) magnesium, (4) tin, (5) lead, instead of iron.   Note carefully, as before, the action in each case.   If necessary, the liquid may be gently heated in order to start an action ; but directly the action starts the tube should be withdrawn from the flame.

iii. **The products of the action.**—Carefully evaporate, nearly to dryness, the liquid in each case in which an action has occurred and allow the liquid to crystallise.   If no crystals are obtained evaporate to dryness.

Action of acids on metals.—When dilute sulphuric acid is added to some iron filings, no action appears to result at first.   After a short time, however, or on warming the tube, an energetic action occurs and the liquid effervesces, owing to the escape of a gas.   This gas does not ignite a glowing splinter, and if a lighted taper or match is placed at the mouth of the tube, the gas ignites with a slight explosion.   The liquid, if filtered from the undissolved iron, has a greenish colour and, on careful evaporation, deposits green crystals, which resemble the green vitriol previously examined (p. 22).

Nitric acid also dissolves iron, but the gas previously obtained will most probably not be given off.   Some yellowish fumes with a choking smell will, however, pro-

bably be evolved and a yellowish liquid will result, which on evaporation leaves behind a yellow coloured solid.

Hydrochloric acid also dissolves iron, an energetic action resulting, whilst a gas is evolved which does not cause a glowing splinter to ignite, but which ignites itself with a slight explosion when a lighted match is applied ; apparently the gas evolved resembles that obtained by the action of sulphuric acid on iron. A yellow liquid also results, which on evaporation leaves a yellow solid as a residue.

Zinc, like iron, is acted on by all three acids. A colourless liquid results in each case, which, on evaporation, leaves either a white solid or colourless crystals. In the case of sulphuric and hydrochloric acids, the gas previously observed appears to be given off, but with nitric acid it is not evolved. The action with nitric acid, however, depends so much on the temperature, and on the strength of the acid, that various reactions may occur. It will probably be found, however, that a gas is evolved in which a lighted match continues to burn, or which may even cause a glowing splint to ignite. A small quantity of yellow fumes may also be obtained.

Copper is not acted on by either dilute sulphuric acid or hydrochloric acid ; dilute nitric acid, however, attacks it very vigorously, and copious fumes of a reddish brown colour and choking smell are evolved. When the copper has entirely disappeared a blue solution remains, which, by evaporation and crystallisation, deposits blue crystals.

Magnesium dissolves very readily in all three acids, forming colourless solutions. With sulphuric acid and hydrochloric acid the combustible gas, previously obtained by the action of zinc on these acids, is again obtained. With nitric acid also, if sufficiently dilute, the same combustible gas results ; but, if the acid is more concentrated, some of the other gases, obtained by the action of nitric acid on the other metals examined, will most probably be produced.

Tin and lead are both unacted upon by the dilute sul-

phuric and hydrochloric acids, but both are attacked by nitric acid, especially on warming. The tin, however, does not dissolve, but forms a white powder, whilst in neither case is the inflammable gas given off.

The action of nitric acid on metals appears to be of a more complicated nature than that of the other two acids, inasmuch as different gases may be evolved—the student will probably have observed three—and, even with the same metal, the products of the action vary with the strength and temperature of the acid. This complication makes nitric acid unsuitable for our study at this stage, and attention may be confined to the action of the other two acids employed.

In every case where the dilute sulphuric, or hydrochloric acid, acted on a metal, a gas was given off which would not cause a glowing splint to ignite, but which ignites itself with a slight explosion. This is the case, for example, when either of the acids acts on iron, zinc, or magnesium. The combustible gas evolved in so many of the cases examined is named **hydrogen**, which we may regard, therefore, as generally formed by the action of an acid upon a metal. This gas must now be prepared and examined.

45. PREPARATION AND PROPERTIES OF HYDROGEN.

i. **Preparation of hydrogen.**—Select a flask, or bottle, and fit it up as is shown in Fig. 43. Be very careful that the stopper and the tubes respectively fit very closely. Into the flask put enough granulated zinc to cover the bottom. Pour some water upon the zinc. Arrange the delivery tube in the trough as you did when you were making oxygen. Pour a little sulphuric acid down the thistle-headed acid funnel, and be quite sure that the end of the funnel dips beneath the liquid in the flask. Do not collect bottles of the gas until you are sure pure hydrogen is being given off, which you can find out in this way. Fill a test-tube with water and invert it over the end of the delivery tube. When it is full of gas, still holding it upside down, take it to a flame (which should not be near the flask you are using) ; notice that there is a slight explosion. Continue this until the hydro-

gen burns quietly down the test-tube.   When this happens you
may proceed to fill one or two bottles.[1]   When the bottles have
been filled, it is better not to remove them from the water until
you want to use them.   Collect also a soda-water bottle half full
of the gas, and fill two test-tubes with the gas
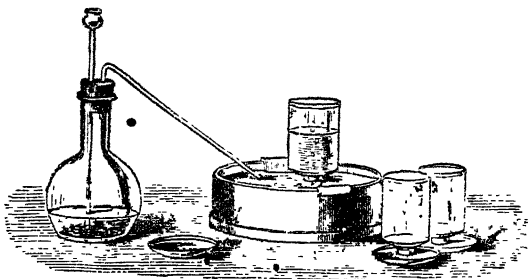
FIG. 43.—Preparation of hydrogen

Caution.—*Be careful not to bring a light near the thistle funnel
or tube delivering the gas, even when the action in the flask
seems to have ceased, or a dangerous explosion may occur.*

*Be careful also that none of the acid used gets upon your
fingers or clothing.*

Do not throw away the acid, etc., left in the flask, but reserve
them for later examination.

ii **Hydrogen escapes from a tube held mouth upwards.**—Take
the two test-tubes of gas ; hold one mouth upwards and the
other mouth downwards for an equal time—say one minute.
Now, apply a lighted taper to both, and observe that a slight
explosion results in the case of that which has been held mouth
downwards, but no trace of hydrogen is found in that which has
been held mouth upwards.   Consider carefully the conclusion to
which this experiment leads (Fig. 44).

iii. **Hydrogen may be poured upwards.**—Take a full jar of the
gas and hold it mouth upwards below a second smaller jar held

[1] In order more speedily to obtain the hydrogen in a state of purity,
raise the thistle funnel until the end is just above the level of the liquid
in the flask or bottle.   Close the end of the delivery tube with your
finger, so that the gas has to escape through the thistle funnel and from
the bottom of the flask.   After waiting for one or two minutes the
thistle funnel should be again pressed down into the liquid and the gas
collected.   It will most probably be found that the gas will fulfil the
test for purity given above.   Be sure that no flame is near the apparatus.

mouth downwards, as shown in Fig. 45. On testing with a lighted taper observe that the gas has left the lower jar, and has filled the upper. Many experiments, as the filling of balloons or
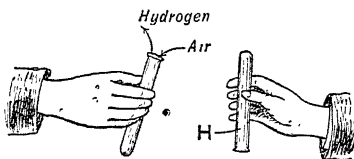
FIG. 44.—Hydrogen escapes from a tube held mouth upwards.
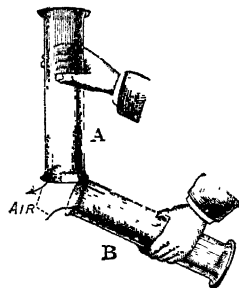
FIG. 45.—Hydrogen can be poured upwards.

soap bubbles, may also be performed to demonstrate the extremely low density of hydrogen.

iv. **Hydrogen burns but extinguishes a flame.**—Test one jar of the gas by means of a lighted match or taper, holding the jar mouth downwards. Observe that the gas burns at the mouth of the jar, but that the taper is extinguished when thrust into the jar; on being taken out again, the taper becomes alight on passing through the flame of the burning hydrogen (Fig. 46).
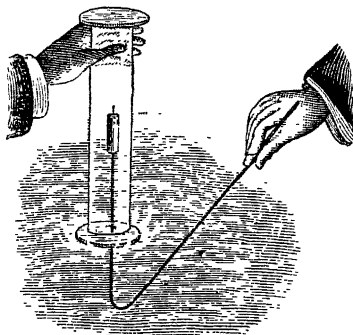
FIG. 46.—Hydrogen burns but does not support combustion.

v. **Hydrogen forms an explosive mixture with air.** —Wrap your hand in a duster, and with it hold the soda-water bottle. Take it out of the water so that the water runs out, and the bottle is now filled with a mixture of hydrogen and air. Apply a light and you will not fail to observe that an explosion results.

N.B.—Be careful to direct the mouth of the bottle away from everybody.

vi. **The flame of burning hydrogen.**—Fit a doubly-bent tube, drawn out to a point, to a hydrogen generator, as shown in Fig. 47. After pouring a little sulphuric acid down the thistle funnel, and proceeding as described in the footnote, Expt. 45, 1., collect a test-tube of the gas issuing from the straight tube, and hold the mouth of the test-tube near a flame, *which must be a few feet away from the generator.* If the gas does
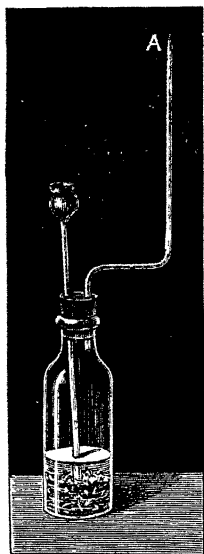
FIG. 47.—Burning of hydrogen.
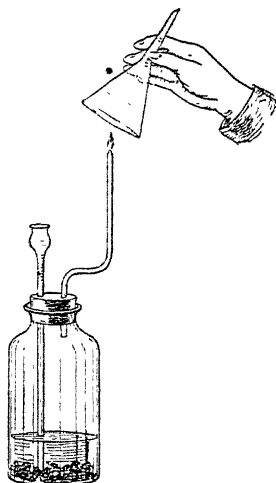
FIG. 48.—Film of liquid formed by burning hydrogen.

not burn quietly, but with a 'pop' or a squeaking noise, collect another tube, and after two or three trials it will burn quietly, with a blue flame. *When you can carry this flame of burning hydrogen, in the test-tube, to the apparatus from which the gas is being produced, do so, and use it to ignite the gas escaping from the pointed tube.* If you remember always to do this, there can be no danger, for when you are able to carry a flame of hydrogen in a test-tube of the gas, for a distance of two or three feet to the generating apparatus, you may be sure that the hydrogen issuing from the apparatus is not mixed with air.

Light the hydrogen in this way.

Observe that it burns with a pale blue flame, which after a time becomes yellow. This coloration is due to the glass becoming hot, and some of the substances in it being burnt in the flame of hydrogen. Hold a funnel over the flame of the burning hydrogen, and observe that it soon becomes covered with a mist, *i.e.* that a thin film of liquid condenses on the funnel (Fig. 48).

vii. **Hydrogen can be collected by upward displacement.**—Substitute for the delivery tube in Expt. 45, i., a tube bent twice at right angles, as shown in Fig. 47, and arrange a jar on a retortstand in an inverted position. Place a test-tube over the upright tube, and allow it to stay there for a minute, and test the gas as described in Expt. 45, vi. When it is pure, substitute the inverted jar, and after a few minutes lift it off the stand and apply a light, first taking the precaution to wrap a duster round the jar, and to hold it away from your face. The jar will be found to contain hydrogen.

viii. **The liquid left in the flask.**—Filter off the liquid in the flask from the undissolved zinc (sufficient zinc should be used to leave a quantity still undissolved ; if all has disappeared add more and wait till the action ceases). Partially evaporate the liquid and allow it to crystallise. You will find that clear, colourless crystals are formed. Examine these crystals and sketch the most perfect. Heat some of the crystals in a tube and observe that they melt, give off water (which can be collected and proved to be water, as in Expt. 11, ii.), and leave a white powder.

Heat this white powder strongly, and notice that it becomes yellow, but again turns white when cold. Consider carefully what other powders behaved similarly (p. 57).

Preparation of hydrogen.—A convenient method for the preparation of hydrogen is to act upon a metal like zinc, with dilute sulphuric acid. Iron could be used, but it is not so clean to work with, and the hydrogen given off is frequently mixed with other gases, owing to impurities in the iron. Magnesium could also be used, but it is much more expensive than zinc. It may be noted here that sulphuric acid has only a very slight action on *perfectly pure* zinc, but it is not likely that the zinc employed will be quite pure. The apparatus suitable for the preparation of hydrogen in this way is described in the experimental work. Owing to

the insolubility of hydrogen in water, it can be collected, in the same way as oxygen, over the pneumatic trough.

Physical properties of hydrogen.—Hydrogen is a gas which possesses no colour, no taste, and no smell. It is very much lighter than air. For this reason soap bubbles, or balloons filled with it, rise in the air when free to move. It has been found by many carefully performed experiments that hydrogen is the lightest substance known.

Hydrogen will burn, but it will not allow things to burn in it.—When a lighted match is brought to the end of a delivery tube, from which pure hydrogen is issuing, the gas catches fire and continues to burn with a flame which is at first almost colourless, but very soon changes to a bright, golden yellow colour; this is due to substances contained in the glass. Hydrogen is therefore said to be *combustible.*

If a jar of hydrogen is held with its mouth downwards, and a lighted match (or a piece of taper) fixed to the end of a wire is pushed up into the jar, the hydrogen, as in other cases, catches fire, and continues to burn at the mouth of the jar; but the flame of the match or taper is put out, showing that things will not burn when surrounded by hydrogen gas. If the match or taper is slowly drawn out of the jar, it will be again ignited as it passes through the hydrogen flame at the mouth of the jar Thus, though hydrogen is a combustible gas, it *will not support combustion* in ordinary circumstances.

Hydrogen being lighter than air can be poured upwards.—When you wish to pour water from one vessel to another, you know very well that the vessel *into* which water is to be poured is placed below, and the vessel *from* which the water is to be poured is placed above and tilted. This is because water is heavier, bulk for bulk, than air. Similarly, if it were required to pour a gas considerably heavier than air from one vessel to another, like the gas which is produced when carbon is burnt in oxygen, you would pro-

ceed in just the same way as with water.   Since, however,
hydrogen is much lighter than air, the order of things must
be exactly reversed.   The vessel *into* which hydrogen is to
be poured must be placed above and be inverted, while the
vessel *from* which the gas is to be poured must be below
and held with its mouth upwards.   This method of pouring
hydrogen from one jar to another is shown in Fig. 45, and
is always used when it is required to pour a gas lighter than .
air from one vessel to another.   For the same reason hydro-
gen escapes very readily from a vessel held mouth upwards,
far more rapidly than from a vessel held mouth down-
wards.   Although hydrogen is lighter than air, it must not
be thought, however, that it could be kept indefinitely in a
vessel held with the open end down.   Even in this case the
hydrogen would *diffuse* and escape into the air.

The reason for the device adopted in Expt. 45 (footnote),
in order to obtain the hydrogen free from air, is now readily
understood.   As the hydrogen is lighter than the air, which the
flask at first contains, it tends to pass directly to the top of
the flask, and if drawn off from there will only very slowly .
drag out with it the air of the flask, which tends to accumu-   ·
late at the lower parts, *i.e.* just above the liquid.   When the
funnel is drawn just above the liquid, however, and the
delivery tube is closed, the gas is drawn off from the lower
portions.   The hydrogen collects at the top of the flask,
forces out the lower layers, which contain chiefly air, through
the thistle funnel, and hence the flask soon contains only   ·
pure hydrogen, or at least hydrogen free from air.

Examination of liquid remaining in flask.—If when the
chemical action in the flask has completely stopped, the
liquid is filtered from the still undissolved zinc, as previously
explained, and then partially evaporated in a basin and
afterwards allowed to crystallise, clear, colourless crystals
are formed, in most cases of a long rod-like form, or if
rapidly crystallised of fine needle-like form.   If heated

in a tube, these crystals melt, give off *water of crystal-lisation*, and leave a white powder. This white powder on strongly heating changes colour, becoming yellow, but again becomes white on cooling. This peculiar change has been observed previously in the case of oxide of zinc. The crystals appear, therefore, to form oxide of zinc on strongly heating. If careful examination and comparison be made of the *strongly* heated crystals and oxide of zinc, this will be found to be really the case. The *crystals contain the zinc*, which could, if necessary, be obtained again from them. The hydrogen has, therefore, most probably been derived from the acid. The other part of the acid has combined with the zinc to form the crystals, which consequently consist of zinc and part of the sulphuric acid, and are known as *zinc sulphate*. We may therefore state *sulphuric acid and zinc form hydrogen and zinc sulphate*. Or, the same fact may be expressed in another way :

	when				
SULPHURIC ACID	acted upon with	ZINC gives	ZINC SULPHATE	and HYDROGEN.	

The action therefore appears to be this, that the zinc takes the place of the hydrogen previously contained in the acid, this hydrogen being set free. This is a very general action ; it occurs with many metals and acids.

QUESTIONS ON CHAP. VIII.

1. By what means is hydrogen most conveniently obtained? What are the principal properties of the gas?

2. Describe the preparation and collection of hydrogen by the action of (*a*) sodium, (*b*) iron, on water. What other products are also formed? How would you obtain each of the products? Describe briefly their various characters.

3. In the preparation of hydrogen from zinc and sulphuric acid, what other product is formed? How would you separate it from the liquid? Describe briefly its appearance.

4. Describe three experiments to show the most characteristic properties of hydrogen.

5. Describe the apparatus you would use for the production and collection of hydrogen gas ; name the materials required and describe the properties of the gas.

6. A jar of oxygen, a jar of nitrogen, and a jar of hydrogen are put before you. How can you tell which is which ?

7. How would you demonstrate (1) that hydrogen is a light gas, (2) that it does not support combustion ?

8. Why is it that hydrogen which burns quietly when ignited at the end of a delivery tube, explodes if mixed with air and then ignited ?

9. What tests would you apply in order to see if hydrogen prepared in the customary method is sufficiently pure to ignite.

CHAPTER IX.

WATER.

46. COMPOUND FORMED BY BURNING HYDROGEN.

i. **Liquid formed by burning hydrogen.**—Arrange a flask or
bottle for the preparation of hydrogen. Also fit to a large
" Winchester Quart" bottle [1] an india-rubber stopper with two
holes. Through one hole pass a short tube bent at right angles,
and through the other pass a long tube, reaching to the bottom
of the bottle, also bent at right angles, and on the outer end
of which is joined some india-
rubber tubing. Fill this bottle
completely with water, insert the
stopper and connect, by india-
rubber tubing, to the delivery tube
of the hydrogen flask, as soon as
pure hydrogen is being evolved.
By this means the water will be
forced out of the long tube and the
Winchester bottle will be com-
pletely filled with hydrogen (Fig.
49). Remove the delivery tube of
the hydrogen apparatus from the
india-rubber connection, and close

FIG. 49.—Filling a 'Winchester'
bottle with hydrogen.

the latter by a spring clip. Now attach to the water-tap the
other end of the india-rubber tube, which is joined to the
long tube. To the india-rubber tube, fitted to the short tube,
affix a U-tube containing either chloride of calcium, or some
pieces of pumice stone soaked in sulphuric acid, either of
which will absorb moisture from, and dry, a gas passing
through it. To the other end of the U-tube connect a
piece of glass tube bent at right angles and drawn out to a

[1] The large bottles holding about 2 quarts in which acids, etc., are sold
are known as Winchester quart bottles, or often briefly "Winchesters."

point. When the water-tap is turned on, the hydrogen will be forced through the drying tube, dried, and can then burn at the end of the glass tube. It should be allowed to burn, as shown in Fig. 50, under a glass retort kept cool by a stream of water flowing in at the tubule and out at the end of the neck. By this means a perfectly regular stream of pure dry hydrogen can be obtained and burnt, and the rate of burning may be readily regulated by means of the tap by which the water flows into the Winchester bottle.
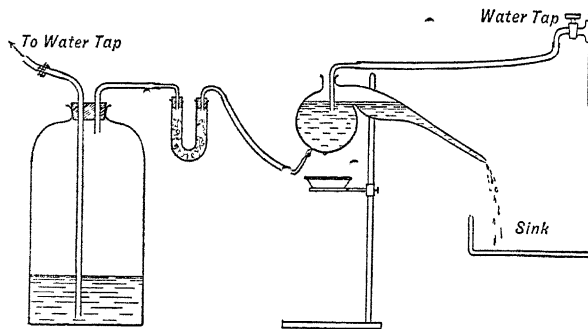
FIG. 50.—Collection of liquid formed by burning hydrogen.

Observe the formation, on the outside of the retort, of a clear liquid, which collects and drops into a clean beaker placed to receive it.

In this way obtain sufficient of the liquid to enable its properties to be examined, using additional Winchester bottles of hydrogen if necessary. Much time is saved if a number of pupils all work at the same experiment, and the liquid formed by all the hydrogen flames is added together.

ii. Nature of liquid formed by burning hydrogen.—Observe that the liquid is colourless, odourless, and tasteless. As already learnt, proceed to find the boiling point, freezing point, and density of the liquid formed when hydrogen burns in the air. The boiling point is found to be 100° C., the freezing point 0° C., and the density 1. The liquid is consequently pure water.

Evaporate some of the liquid in a basin, and notice that no residue is left.

Test some of the liquid with litmus paper. Observe that it is neutral.

The burning of hydrogen.—When an apparatus is ar-

ranged as in Fig. 50, so that a hydrogen flame burns below a surface which is kept cold, it is found that a liquid is continuously formed, which, as the experiment proceeds, collects drop by drop in a beaker put to catch it. Since this liquid is formed by the burning of a perfectly dry gas, it must be a substance which hydrogen forms with something out of the air. To be sure that dry hydrogen is burnt, the gas is first passed through a tube containing lumps of calcium chloride, a substance which has the power of absorbing moisture.

The liquid formed as hydrogen burns is water.—When the liquid produced by burning dry hydrogen in air is examined, it is found to have all the properties of pure water. Thus, its boiling point is 100° C., its freezing point is 0° C., and its density is 1 ; when evaporated to dryness, no residue is left behind, and it does not change the colour of litmus paper used for testing it.

These are all characteristics of pure water, and it is therefore quite certain that the liquid, which is formed as hydrogen burns in air, is water.

Water is oxide of hydrogen.—We have already learnt that substances in burning unite with the oxygen of the air to form oxides. Unless hydrogen is quite exceptional, therefore, it follows that the burning of hydrogen is also its union with the oxygen of the air to form an oxide, and that this oxide is water. *Water* is therefore the *oxide of hydrogen*. That hydrogen is not exceptional may be readily proved by lighting a jet of hydrogen and placing around it (*a*) a jar of nitrogen, (*b*) a jar of oxygen. The burning hydrogen, like a burning taper, etc., is extinguished in the nitrogen, but continues to burn in the oxygen, showing that the hydrogen resembles in this respect other combustible substances.

If water is an oxide of hydrogen, it might be possible to obtain the hydrogen from water by suitable means. To do this we must act upon it with some substance which will abstract from it the oxygen. We have found that iron does

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not rust in, *i.e.* abstract oxygen from, pure water, but it is quite possible that it might do so if heated strongly, or that other metals might do so either from cold or heated water.

47. TO OBTAIN HYDROGEN FROM WATER.

i. **Action of heated iron on water.**—Place some iron filings in the hard glass tube CA, or better in a piece of ordinary iron gas pipe, and let the end A dip under water. (If an iron tube is
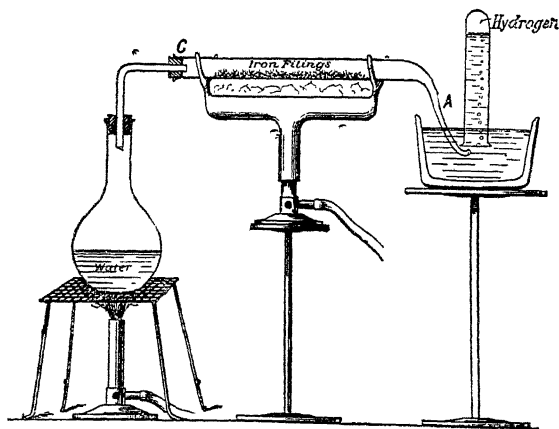
FIG. 51.—Action of heated iron on steam.

used a delivery tube dipping under the water must be fitted to its end by means of a bored cork.) To the end C fit a delivery tube from a flask, or tin can, containing water. Heat the iron filings well, and boil the water in the flask so that steam passes over the heated iron, and then into the water, where it condenses. After boiling for sufficient time for the air to be driven out from the flask and tubes, place an inverted test-tube of water over the end A, and note that the steam is not completely condensed, but that minute bubbles ascend to the top of the test-tube. When you have obtained by this means a sufficient quantity of gas (half a small test-tube), first disconnect C from the flask of water, then stop the boiling. Close the end of the test-tube with your thumb, and holding a lighted match to the mouth, open the tube. Observe that the gas *burns* with

the blue flame characteristic of hydrogen Examine the iron filings in the tube *C-A*, and see that a quantity of rust, *i.e.* oxide of iron, has been formed (Fig. 51).

ii. **Action of sodium on water.**—-Place a small piece of sodium[1] in water in an evaporating basin, and quickly put a large glass shade over the latter ; observe the action. Feel the water left after the sodium has all disappeared, and test it with red litmus. Evaporate away the water. Note the residue Compare the properties of the liquid with those of the solution of oxide of sodium.

iii. **Collection of the gas which sodium displaces from water.**—Place a small piece of sodium in a small piece of lead tubing, the ends of which are nearly closed, and gently drop the lead into a pan of water. A gas is seen to come off (Fig. 52). Collect this in an inverted tube full of water, and by this means obtain three test tubes of the gas. Observe that the gas is colourless and odourless. Test the gas as in Expt. 45, and observe it is lighter than air, inflammable, and that mixed with air it is explosive, thus confirming the view that it is hydrogen.
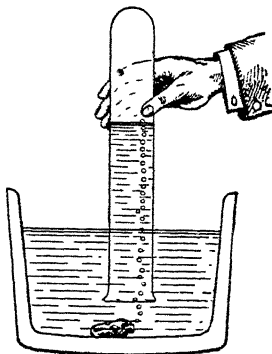
FIG. 52.—Action of sodium on water.

iv. **Action of magnesium on water.**—Place a small piece of ordinary magnesium ribbon in a test-tube completely filled with water and inverted over a basin of water. Observe that minute bubbles form on the magnesium, even though attempts to completely get rid of any air bubbles adhering to it be made before it is placed in the test-tube. Note that these minute bubbles *slowly* collect at the top of the tube. Place on one side, and when sufficient quantity of gas has collected (this may take some weeks) test the gas by a lighted match, and see that it behaves like hydrogen.

v. **Action of magnesium on steam.**—Arrange apparatus as in Expt. 47, i., or as shown in Fig. 53, the end *d* being fitted with a delivery tube dipping under water, but in the glass tube place magnesium ribbon in place of the iron filings. Heat the magnesium strongly in the current of steam, and observe that it becomes converted to a white powder—it may even burn—and that hydrogen collects in the test-tube.

[1] Read again previous warning regarding the use of sodium (p. 84).

Action of heated iron upon water in the form of steam.
—To make as large a surface of heated iron as possible
come into contact with the water, it is best to use iron
filings in studying the action of steam upon hot iron. The
filings are contained in a hard glass or iron tube, which
will not melt when strongly heated. The tube is fitted
at each end with india-rubber stoppers, through each
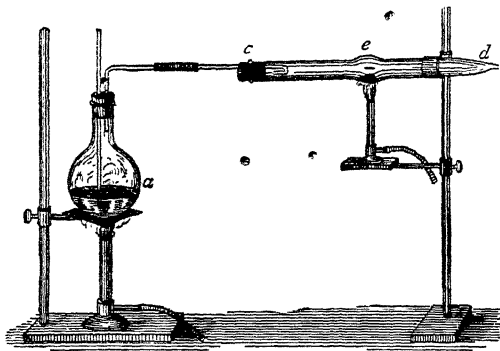
FIG. 53.—Action of magnesium on steam.

of which a hole is made to take a glass tube. One of these
tubes is connected with a flask, or other vessel, in which
water may be boiled. The other tube is bent to make a
delivery tube, which dips into a trough of water.

The iron filings in the tube are strongly heated, and
the water in the flask is made to boil vigorously. At first,
air escapes through the water in the trough; and, after
that, bubbles of steam pass from the end of the delivery
tube and are condensed by the water in the trough. When
this occurs, a tube full of water is inverted over the end of
the delivery tube, and, as it is watched, it is noticed that
small bubbles of gas pass up into the tube and collect there.
When a lighted match is brought in contact with the gas
which collects in the test-tube, the gas is found to burn in

the manner characteristic of hydrogen. Moreover, an examination of the iron filings shows that they have been partly changed into red rust, *i.e.* oxide of iron.

The iron has, therefore, taken the oxygen from the water forming oxide of iron, the hydrogen being liberated. We may state that under these conditions

IRON and WATER (steam) form OXIDE of IRON and HYDROGEN.

Other metals besides iron will displace hydrogen from water.—It has been seen that before iron can set hydrogen free from water it must be heated. Some other metals will do the same thing without being heated. It has been found that a metal called *sodium* will displace the hydrogen from water even when it is cold.

The action of sodium on water.—Though sodium is heavier than naphtha (the liquid in which it is always kept), it is lighter than water. Consequently, when a piece of sodium is put upon water, it floats. When it touches the water, however irregular its shape, the piece of sodium almost immediately becomes quite round, and darts about from one place to another, spinning the whole time. All the while it gets smaller and smaller, and just before its disappearance there is often a little explosion, and small pieces of sodium are scattered in every direction. The greatest care must be taken that none of these tiny fragments hit the face.

Collection of hydrogen displaced by sodium.—If a piece of sodium is put into a weighted roll of wire-gauze, or in the lead tube described, which is dropped into water, then, as soon as the sodium enters the water, bubbles of gas begin to appear. These can easily be collected by holding an inverted test-tube full of water over the tube or roll of gauze. In this way, tubes can be filled with the gas which sodium displaces from the water, and it may be readily seen that

the gas is really hydrogen, being inflammable, forming an explosive mixture with air, and being very much lighter than air. The liquid left, after the action of the sodium, is alkaline, and has a soapy feel. It will be recognised as that previously obtained by dissolving the *oxide of sodium*.

Action of magnesium.—Magnesium ribbon will also displace the hydrogen from cold water, but it does so very slowly. If, however, magnesium instead of iron be heated in a tube as in Expt. 47, and steam passed over it, it will be found that the magnesium will rapidly unite with the oxygen and liberate the hydrogen. The magnesium will even burn brightly in the current of steam. It is clear that all these experiments confirm the view that water is the oxide of hydrogen.

Quantitative examination of the action of acids on zinc or magnesium.—It has been found that zinc, or magnesium, has the power of displacing hydrogen from sulphuric or hydrochloric acid. The question may be asked, does the zinc displace its own weight of hydrogen, or is there *any* connection between the weight of the zinc, or magnesium, and the weight of the hydrogen it displaces? To answer these questions the following experiments may be performed:

48. WEIGHT OF HYDROGEN DISPLACED FROM ACIDS BY ZINC OR MAGNESIUM.

i. **Weight of hydrogen displaced from sulphuric acid by 100 gms. of zinc.**—Weigh out a few pieces of zinc in the form of thin sheet ; 4 to 6 grams is a convenient weight. Place this zinc in a flask containing water and closed by an india-rubber stopper with two holes, through which pass, (1) a tube reaching below the level of the water in the flask, and closed outside by a piece of india-rubber tubing, into which is inserted a piece of glass rod. (2) A tube packed with cotton wool, or better with glass wool and calcium chloride (*a* and *b* in Fig. 54). Obtain a small tube and place in it some strong sulphuric acid. By means of a string tied round the tube lower it into the flask, and place in the india-rubber stopper so that the tube is held in position and no acid is spilt. Now very carefully weigh the whole flask with

the acid, water, and zinc. Remove the stopper and gently lower
the tube so that it rests inside without spilling the acid ; place the
string inside and tightly re-insert the stopper. Now gently
tilt the flask until the acid runs out of the tube and acts upon
the zinc. Hydrogen is evolved and escapes through the tube (b)
in which it is dried ; so that only dry hydrogen escapes.
When all the zinc has dissolved and the action has ceased,
remove the indiarubber tube from the end of the tube (a), and
draw a current of air through the flask, so that all the hydrogen
in it shall be displaced by air. Again weigh the flask ; the
apparatus will be found to have decreased in weight. The
decrease must, obviously, be the weight of the hydrogen which
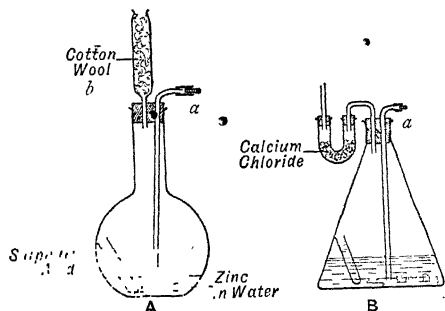
FIG. 54.—Weight of hydrogen displaced from acid.

has escaped from the flask. Calculate the weight which would
have been evolved by the action of 100 gms. of zinc. Enter
your results as follows :

Weight of zinc = 5·46 gms.

Weight of flask, zinc, and acid (before action) = 79·894 gms.
Weight of flask,   ,,   (after action) = 79·728 gms.

Weight of hydrogen evolved,   -   -   0·166 gm.

5·46 gms. of zinc displace 0·166 gm. of hydrogen.

100 gms. of zinc   ,, $\dfrac{0·166 \times 100}{5·46}$ gms. of hydrogen.

= 3·05 gms. nearly.

ii. Repeat the experiment, using hydrochloric[1] in place of
sulphuric acid.

[1] If a number of students are working together it may suffice for some
to employ sulphuric acid whilst others employ hydrochloric acid.

iii. **Weight of hydrogen displaced from acids by magnesium.—**
Repeat, employing magnesium instead of zinc with both acids—a smaller weight, 1 to 2 grams, will be sufficient. In each case enter your results as previously shown.

Quantitative nature of action of acids on metals.—It will be found by the experiments just described that the quantity of hydrogen displaced by the zinc is independent of the acid used, and that whichever acid is taken, 100 grams of zinc always displace about 3·1 grams of hydrogen. The values found may vary slightly ; but, although the experiment is one which requires considerable care and accuracy in weighing, yet the values will probably not differ much from that given, which is the value always obtained when the experiment is performed with complete precautions for the drying of the hydrogen, etc. It follows that the weight of zinc able to displace 1 gram of hydrogen is about 100/3·1, or 32·5 grams We may say, then, that 32·5 grams of zinc are equivalent to one gram of hydrogen in this respect, that this weight of zinc can displace one gram of hydrogen from acids. It will similarly be found that 100 gms. of magnesium always displace 8·3 gms. of hydrogen ; the weight of magnesium, therefore, which will displace one gram of hydrogen, or which is *equivalent* to one gram of hydrogen is 100/8·3 = 12 gms. ; 12 is hence said to be *equivalent weight* of magnesium. The volume of hydrogen displaced by these metals may also be found.

49. VOLUME OF HYDROGEN EVOLVED BY THE ACTION OF ZINC, OR MAGNESIUM, ON ACIDS.

i. **Volume of hydrogen displaced by 100 gms. of magnesium.—**
Weigh out about 0·25 gms. of magnesium (do no employ more than this) and roll it into a small coil. In a graduated 250 c.c. jar place some dilute acid—hydrochloric acid preferably, although sulphuric acid will do. Invert this in a large trough of water, and quickly insert the coil of magnesium under the mouth of the jar. Observe the rapid formation of hydrogen, the bubbles of

which, by clinging to the magnesium. cause it to rise to the surface, which descends rapidly as the gas collects above (Fig. 55). When all the magnesium has dissolved, allow the jar to cool, and measure the volume of gas in the cylinder. It is best to lower the cylinder in the trough until the level of water inside and outside is the same, so that the pressure of the gas is equal to that of the air. Next calculate the volume which would have been evolved by the action of 100 grams of magnesium. Observe the temperature of the laboratory, and the height of the barometer, and, in recording your result, state also the temperature and pressure at which this volume of gas was measured. Record your results thus :

Weight of magnesium $= 0.246$ gm.

Volume of hydrogen evolved $= 243$ c c.

Temperature 17°C. Height of barometer 752 mms.

0.246 gm. magnesium displaces 243 c.c. hydrogen at 17° and 752 mms.

100 gms. magnesium displace $\dfrac{243}{0.246} \times 100$ c.c. hydrogen at 17° and

752 mms.

$= 98,800$ c.c. or 98.8 litres.

At 0° and 760 mms. this volume would be :

$$\frac{98.8}{1} \times \frac{273}{290} \times \frac{752}{760} \text{ litres.}$$

$$= 92.3 \text{ litres.}$$
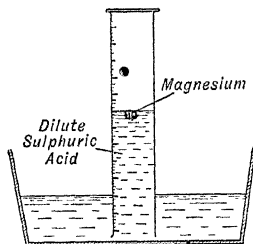
FIG. 55.—Volume of hydrogen displaced by magnesium
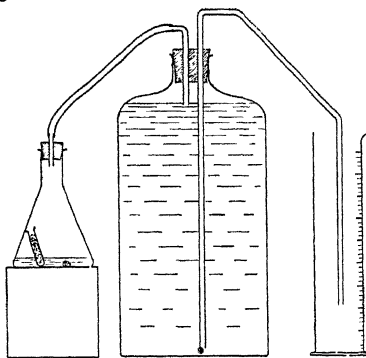
FIG. 56.—Volume of hydrogen displaced by zinc from acids.

ii. Volume of hydrogen displaced by 100 gms. of zinc.—This volume may be determined in a manner similar to that employed

for the magnesium. Or, it may be conveniently found by employing a method analogous to that used to find the volume of oxygen evolved from 100 gms. of potassium chlorate (Fig. 42). Place some zinc, previously weighed (about 0·7 gm.) in a small flask containing water and fitted with an india-rubber stopper, through which passes a glass tube, fitted, as shown in Fig. 56, to the Winchester bottle filled with water. Carefully place in the flask a small test-tube containing concentrated sulphuric acid, taking care not to spill the acid. (This is best done by previously tying a piece of string round the tube and lowering it by means of this string into the flask.) Replace the india- · rubber stopper and tilt the flask so that the acid runs out of the tube. Hydrogen will be evolved and will displace the water. The results may be calculated as in the case of the magnesium.

Density of hydrogen.—The volume of hydrogen evolved by the action of 100 gms. of zinc, or magnesium, on acids may be found in the manner described in Expt. 49, i. and ii. It will be found that by the action of 100 gms. of magnesium about 93 litres of hydrogen, and by the action of 100 gms. of zinc about 34·5 litres of hydrogen are evolved. As we have already found the weight of hydrogen evolved by the action of these metals, it is clear that we have the connection between the weight and volume of hydrogen. Thus, the experiments with magnesium prove that 93 litres of hydrogen weigh 8·3 gms., whilst those with zinc prove that 34·5 litres of hydrogen weigh 3·1 gms. Each of these gives the weight of 1 litre of hydrogen as 0·09 gm., and the concordance of the results may be taken as an indication of their validity.

This weight is important, and must be remembered. We may write, therefore,

                1 litre of hydrogen weighs 0·09 gms.

                1 c.c. of hydrogen weighs 0·00009 gms.

Density of other gases compared with hydrogen.—Hydrogen being the lightest gas known the densities of other gases are frequently compared with that of hydrogen, and this relative density is called the **vapour density** of the gas. Thus, the weight of 1 c.c. of oxygen has been found to

be 0.00144 grams. Its density, therefore, compared with
that of hydrogen is $\dfrac{0.00144}{0.00009} = 16$.

This relative density is also frequently called 'the **density**
of the gas. The student must always endeavour to consider,
when the density of a gas is spoken of, whether the absolute
density, *i.e.* weight of 1 c.c., or this relative density is to be
understood.

QUESTIONS ON CHAP. IX.

1. Give diagrams and a short description of an experiment to
prove that water is produced by the combustion of hydrogen.

2. What tests would you apply in order to determine whether a
given colourless, odourless, tasteless liquid was water or not?

3. Write down what you consider to be the physical and chemical
characteristics of water, that is, the properties which are possessed by
water, but by no other substance.

4. Describe a method of liberating hydrogen from water (*a*) at
a red heat ; (*b*) at ordinary temperatures.

5. A current of steam is passed through a tube which contains
pieces of iron heated to redness. What chemical change goes on
within the tube, and what tests would you apply to identify the
gas which escapes from the end of the tube?

6. How would you prove that air contains one of the constituents of
water?

7. Describe carefully a method by which the weight of hydrogen
displaced from acid by 1 gram of zinc can be determined, adding
numbers to indicate the mode of calculation.

8. What chemical reaction takes place when a small piece of
metallic sodium is thrown into water?

9. How would you test whether the solution left is acid or alkaline?

10. By what experiments have you determined the density of
hydrogen?

11. What do you understand by the equivalent weight of a substance?
It was found that 0.36 gramme of magnesium liberated from dilute
sulphuric acid 336 c.c. of hydrogen at N.T.P. Can you deduce the
equivalent weight of magnesium? What further experiments would
you make with this metal to find the equivalent weight of oxygen? (1
litre of hydrogen at N.T.P. weighs 0.0899 gramme.)

12. Sketch the apparatus you would employ, and mention the
precautions you would take, in showing the combustion of hydrogen
in air. By what physical and chemical properties would you prove
that the product is water?

CHAPTER X.

COMPOSITION OF WATER.

50. QUANTITATIVE COMPOSITION OF WATER.

Warning.—The student must be warned that this experiment is dangerous unless care is taken, as any escape or leakage may lead to a serious explosion. **With care** it may be safely done.

1. **To determine the weights of oxygen and hydrogen which combine to form water.**—Obtain a piece of hard glass-tube drawn out at one end; pack in this tube some asbestos, and after this some granulated copper oxide, and finally some more asbestos. See
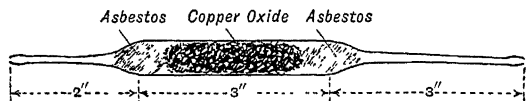
FIG. 57.—To illustrate Expt. 50.

that a current of air can be drawn through the tube, but that the asbestos is packed sufficiently tightly to prevent the escape of any particles of copper oxide. The tube should be now drawn out at the other end, so as to take the form shown in Fig. 57. (The dimensions given will be found convenient.) Obtain a U-tube, fitted with a small bulb, as shown in Fig. 58, and a cork, which will fit into this bulb. Bore a small hole in the cork and fit it on the end of the copper oxide tube. Obtain also a second U-tube. Heat some sulphuric acid and small lumps of pumice in a crucible for a short time, until white fumes begin to be evolved, and then place the pumice and a convenient quantity of sulphuric acid in the U-tubes (Fig. 58). One of these U-shaped tubes serves to dry the hydrogen, in the second the water produced in the experiment is collected. Fit up the apparatus as shown, using the Winchester bottle to contain the

hydrogen, as in Expt. 46, i.   It is important for the copper
oxide tube to slope as shown.   Before filling the Winchester
bottle with hydrogen, allow it to be filled with ordinary air,
and, heating the copper oxide tube by a Bunsen burner flame,
turn on the tap and allow a current of dry air to pass through
the apparatus for a few minutes.   If any trace of moisture is
seen in the copper oxide tube it should be taken off, heated,
and a current of air blown through by a bellows until no trace
of moisture is observed.   The passing of the current of air
will also show whether your apparatus is in good order or not,
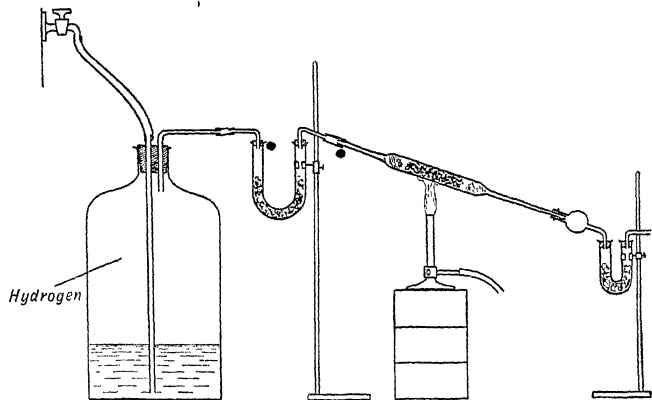
FIG. 58.—Determination of the composition of water.

as bubbles of gas should pass through the acid in both U-tubes.
If satisfactory, the copper oxide tube and the second U-tube
(right-hand side of the diagram) are both carefully weighed, the
cork being kept on the end of the copper oxide tube.   (To
weigh these tubes it is most convenient to suspend them from
the balance by wire.)   Fill the Winchester bottle with *pure*
hydrogen, connect again as in Fig. 58, and, by turning on the
tap slightly, allow a slow current of hydrogen to pass through
the apparatus.   If the apparatus is seen to be completely air
tight (bubbles passing at the same rate through both U-tubes)
heat the copper oxide tube strongly.   You will probably notice
a bright glow in the tube, owing to the combination of the
hydrogen and the oxygen.   Observe also the water which
collects in the bulb of the U-tube and in the lower part of the
copper oxide tube.   When a sufficient quantity of water has
collected and about three-fourths of the hydrogen has been

used, the lower end of the copper oxide tube should be carefully heated to drive off any drops of water which may have collected there. When the bottle is nearly full of water stop the tap and allow the apparatus to cool, without disconnecting, so that no oxygen or air gets into the hot copper oxide tube. When cool, disconnect and again weigh the U-tube and copper oxide tube, noting how, in the latter, the black copper oxide has been converted into bright copper. Enter your results thus :—

Wt. of U-tube after experiment, -	-	36·473 grams
Wt. of U-tube before experiment,	-	35·821 grams
Wt. of water produced, -	-	0·652 gram

Wt. of tube of copper oxide before experiment,	35·562 grams
Wt. of tube of copper oxide after experiment,	34·982 grams
Wt. of oxygen, - - - - -	0·580 grams

The weight of oxygen is 0·580 gram, and that of the water produced from it is 0·652 gram; therefore the weight of hydrogen must be 0·652 − 0·580 gram, that is, 0·072 gram. Write, therefore :—

Wt. of water, - - -	0·652 gram
Wt. of oxygen, - - -	0·580 gram
Wt. of hydrogen, - -	0·072 gram

Therefore 0·072 gram hydrogen combines with 0·580 gram of oxygen, and hence 1 gram hydrogen combines with $\frac{·580}{·072} = 8$ grams of oxygen.

Composition of water by weight.—We must now endeavour to find the composition of water by weight, that is the weights of oxygen and hydrogen which combine to form water. To do this, it should be noticed that we only require the weights of two out of the three substances concerned, *i.e.* if we know the weights of hydrogen and water (or of oxygen and water), the weight of the oxygen (or hydrogen) is readily calculated. The experiment is performed by finding the weights of the oxygen and water, and for this it is

best to use not oxygen itself, but some oxide which readily gives up its oxygen to the hydrogen, so that by weighing the oxide before, and after the experiment, we can ascertain the weight of oxygen which it has lost. The oxide used for this purpose is oxide of copper, a black powder.

The experiment is one of considerable difficulty. *Great care* must be taken if satisfactory results are to be obtained, but it is possible to obtain such results.

If the experiment is performed in the manner described above, it will probably be found that the ratio of the weight of oxygen to that of hydrogen is between $7:1$ and $9:1$. But, if a number of experiments is performed, and the *total weights* of the oxygen and the hydrogen are considered, a ratio of about $8:1$ will almost certainly be obtained. This is the result which has been always obtained where more precautions have been taken, and more elaborate apparatus used than is possible at this stage. We may, therefore, state that *1 gram of hydrogen combines with 8 grams of oxygen to form 9 grams of water*, or that *water is formed of eight-ninths its weight of oxygen combined with one-ninth its weight of hydrogen.*

Proportions of oxygen and hydrogen by volume in water.—We have found the proportions by weight in which the oxygen and hydrogen combine during the formation of water. It is possible also to find the volumes of the gases which unite to form water. For this it is necessary to measure out definite volumes of oxygen and hydrogen, cause them to combine, then measure the volume of gas which remains uncombined and ascertain which gas it is. This is usually done in a piece of apparatus known as an **Eudiometer** (Fig. 59). In its simplest form this consists of a long glass tube closed at one end and graduated in equal volumes, usually cubic centimetres, by divisions marked on the glass. Through opposite sides of the tube, at the closed end, pieces of platinum wire are passed, and fused into the

glass, being so arranged that they do not quite touch one another. Outside the tube the platinum wires are bent into loops to which wires from an induction coil may be attached.
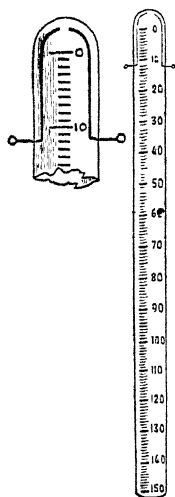
FIG. 59.—A simple form of eudiometer. The arrangement of the platinum wires is shown in the enlarged top of the eudiometer.

Composition of water by volume.— To use the eudiometer it is first completely filled with mercury and inverted over more mercury contained in a trough. A suitable quantity of pure dry oxygen is then bubbled into the tube and the volume (after the necessary corrections for temperature and pressure) is recorded. Pure dry hydrogen is next bubbled into the tube, using a considerable excess of one or other of the gases. The volume is again recorded (with necessary corrections as before), and then, keeping the eudiometer firmly pressed upon a sheet of india-rubber, or felt, at the bottom of the trough, the gases are made to combine by causing an electric spark to pass between the platinum wires inside the tube. As soon as the spark passes, the two gases combine, with a flash of light. The eudiometer is slightly raised from the india-rubber (but of course not above the mercury in the trough), and it is seen that the volume of the gas in the eudiometer is less than before the explosion, and that there is a film of moisture upon the interior of the tube. The corrected volume is again recorded, and the nature of the gas ascertained. We then find the volumes of the gases which have combined, in the manner indicated below :

Corrected volume of oxygen,   -   -   - 12 c.c.
Corrected volume of mixed gases,   -   - 50 c.c.

Therefore, corrected volume of hydrogen = 38 c.c.

Corrected volume after explosion - = 14 c.c.

Gas left ascertained to be hydrogen.

Hence the 12 c.c. of oxygen united with (38 – 14), *i.e.* with 24 c.c. of hydrogen, and we find this result always obtains, namely, *2 volumes of hydrogen combine with 1 volume of oxygen to form water.*

Such a process as this, the formation of a compound from elements, or from simpler materials, is known as a **synthesis**.

We might also find the required ratio by the **analysis** of water, that is by breaking it up into its components, which we can do by passing an electric current through it.

Analysis of water.—This may be done by means of an electric battery, for generating the electric current, and a **Voltameter.** The latter is most simply made by closing the bottom of a funnel by means of a tightly fitting cork through which pass two platinum wires with small plates of platinum attached to the ends remaining in the funnel (Fig. 60). Over these plates are supported two glass test-tubes, and the tubes and part of the funnel are filled
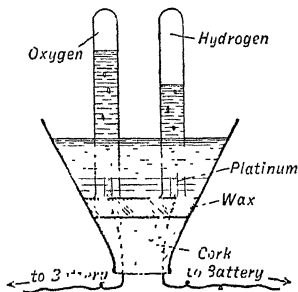
Fig. 60.—A voltameter.

with water to which has been added a little sulphuric acid, as otherwise it offers great resistance to the electric current. The wires from a battery of three or four cells are connected with the ends of the platinum wires, and as soon as the connection with the battery is complete, provided there is clean metal at every junction, bubbles of gas are seen to rise from each platinum plate, and to ascend into the tube and displace the contained liquid. After the experiment has gone on for half-an-hour, the gases may be

J.C.   I   •

tested and their volumes measured. It will be found that the volume of one gas is double that of the other, and that the gas of which there is the larger amount is *hydrogen*, while the other is *oxygen*.

Calculation of composition of water by volume.— If, however, eudiometers and voltameters are not available, the volume ratios can be calculated from the weights and the densities of the combining gases. We have found that 1 gram of hydrogen combines with 8 grams of oxygen. But, one gram of hydrogen occupies 16 times the volume of 1 gram of oxygen (p. 122), and therefore *twice* the volume of 8 grams of oxygen. The volume of hydrogen, therefore, is twice that of the oxygen.

If, therefore, the result has actually been found experimentally it serves to confirm the previous experiment which led to the same conclusion.

We have now found

2 volumes of hydrogen unite with 1 volume of oxygen to from water.

also 1 gram of hydrogen unites with 8 grams of oxygen to form 9 grams of water; and

the density of oxygen is 16 times that of hydrogen.

QUESTIONS ON CHAP. X.

1. What is the composition of water? Describe experiments by which its composition by volume may be found.

2. If twenty volumes of hydrogen and twenty-five volumes of oxygen be mixed in an eudiometer and exploded, what volume of gas would be left, and how would you show which gas it was?

3. What is the action of hydrogen on heated copper oxide? How can this action be used to find the composition of water by weight?

4. What is an eudiometer? Give a sketch of one and explain for what purposes it may be used.

5. How would you prove that when hydrogen is passed over heated copper oxide something is taken away from the latter compound?

6. You are required to prove experimentally that, in the previous question, the part abstracted is really oxygen. How would you endeavour to do so?

7. Describe experiments (1) to show that water may be split up into oxygen and hydrogen, and (2) to show that by the combination of these gases water is produced.

8. What volume of hydrogen would be required in order to unite with 7·3 gms. of oxygen?

9. It is found when 7·26 gms. of an oxide are heated in a current of hydrogen that 6·73 gms. of metal are left. Calculate the percentage composition of the oxide.

10. What (1) weight, (2) volume of hydrogen would have been used in the experiment of Quest. 9?

CHAPTER XI.

EXAMINATION OF CHALK.

51. CHALK.

i. **Examination of chalk.**—Again examine chalk according to the scheme of Chap. II., or read thoroughly your notes on its examination.

ii. **Change produced by heating chalk.**—Place a little powdered chalk (*not* blackboard chalk) on a piece of platinum foil, and heat it strongly for some minutes in the flame of a laboratory burner. If platinum foil is not at hand, heat a lump of chalk on a piece of coarse wire gauze for some time. After heating, shake the powder on to a damp, red litmus paper. Observe that the red litmus paper is, in places, changed to a blue colour.

iii. **Action of acids of chalk**—(*a*) **Hydrochloric acid.**—Place a few pieces of chalk in a test-tube and add some dilute hydrochloric acid. Observe that brisk effervescence occurs, indicating the evolution of gas. Test the gas by a glowing splint of wood and by a lighted match : observe that the lighted match is extinguished. If all the chalk has not dissolved add a little more acid. When all has disappeared evaporate the solution to dryness, and notice the white solid which results. To some of this white solid add a *few* drops of water, and notice the extreme solubility of the solid. Place the remainder of the residue in an evaporating basin and leave it for a day, after which, examine the basin and notice that the solid has liquefied, having absorbed moisture from the air and dissolved in this moisture. It is, therefore, said to be *deliquescent*.

(*b*) **Nitric acid.**—Repeat the previous experiment, using nitric acid instead of hydrochloric acid. You will observe a similar brisk effervescence, and the evolution of the same gas, whilst, as before, a white soluble solid results.

(*c*) **Sulphuric acid.**—Repeat the experiment, this time using sulphuric acid. Notice that the effervescence is not as brisk as in the previous case, and that the chalk does not dissolve, being

apparently unaltered. Test the gas as before. To see that the chalk is really changed, place a little powdered chalk in dilute sulphuric acid in a test-tube and after effervescence has ceased pour off the acid and to the powder add a little hydrochloric acid. You will notice but little if any effervescence.

iv. **Examination of marble.**—Examine marble as above also, and observe that it appears identical with chalk.

Action of heat and acids on chalk.—The more obvious properties of chalk were described on page 18. It was there pointed out that although heat has apparently no action on chalk, yet some action must take place, inasmuch as *lime* is obtained by heating chalk in lime kilns. It is easy to prove by putting some powdered chalk upon a piece of moist, red litmus paper that this substance is unable to change the colour of the paper. If, however, some powdered chalk be strongly heated on a piece of platinum foil in a laboratory burner, and then placed on a piece of moist, red litmus paper, the red colour is changed to blue. The chalk undergoes some change when heated, or it would not acquire this new property.

Acids act on chalk, causing the evolution of a gas which does not support combustion and will not burn itself. This gas may for the present be called *chalk gas*. The action of hydrochloric acid, or of nitric acid, is very vigorous, but that of sulphuric acid is less so. The product which results is, in each case, a white solid, but that produced by the sulphuric acid is only slightly soluble in water, whereas those produced by the other two acids are very soluble. Indeed, the solid resulting from the action of hydrochloric acid on chalk is deliquescent and absorbs moisture from the air—owing to this it has been previously used for drying gas under the name of *calcium chloride*. It is owing to the insolubility of the product formed when sulphuric acid acts on chalk that the action appears so much less vigorous, since the chalk soon gets coated with an outside layer of this product which protects the inner portions.

Examination of marble appears to indicate that this substance, although much harder and denser is very similar to, if not actually identical with, chalk.   Later experiments will prove that the two are really identical in composition.

52. LIME.

i. **Examination of lime.**—Examine lime carefully according to the scheme of Chap. II.   Test it with litmus paper and observe its alkaline nature.

ii. **Action of water on quicklime.**—Add a few drops of water to fresh lime ; observe that the mass gets very hot, swells up, and forms a white powder.   Examine this powder and note that it still preserves the ordinary properties of lime, as indicated in Expts. 52, iii. and iv.

iii. **Action of heat on lime.**—Heat a piece of lime very strongly by means of a blow-pipe flame.   Observe that it glows very brightly, but that no other change is produced.

iv. **Action of acids on lime.**—Examine the action of the three acids on lime in precisely the same way as you examined their action on chalk, obtaining also, by evaporation, the products, etc. (Expt. 51, iii.).

Properties of lime.—Lime, produced by heating chalk or limestone, is a white solid, which, if heated sufficiently, glows and emits a brilliant white light.   It is, on this account, employed for the production of the limelight, where a small, hard cylinder of lime is strongly heated in an oxy-hydrogen or oxy-coal-gas flame.

When water is added to freshly-burnt lime, or *quicklime*, as it is termed, the water combines with it with the evolution of a large amount of heat, sufficient to boil the water if the quantity of lime is large.   This can be seen at any time when bricklayers are preparing lime for making mortar. This addition of water to lime is called *slaking* it, and the altered lime is known as *slaked lime*.   Lime dissolves to a slight extent in water, forming lime-water.

Lime dissolves in acids, but without effervescence; it yields, however, *solid products identical with those obtained from*

chalk. Thus, with hydrochloric acid, the very deliquescent calcium chloride results, whilst with sulphuric acid the but slightly soluble white powder—called calcium sulphate—is obtained ; so that, as in the case of chalk, unless the acid is very dilute and but little lime be taken, it will not dissolve and there will apparently be no action.

53. PREPARATION AND EXAMINATION OF CHALK GAS.

i. **Preparation of chalk gas.**—Into a flask or bottle, fitted like that in Fig. 61, place some chalk. Place the delivery tube in a
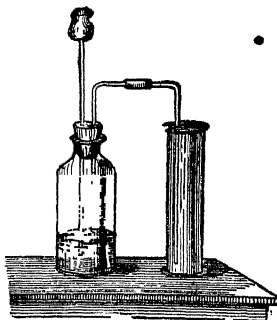
FIG. 61.—Preparation of chalk gas.   FIG. 62.—Pouring chalk gas downwards.
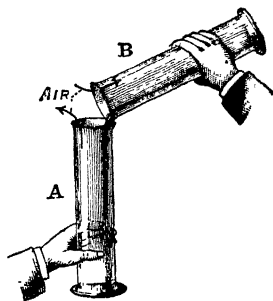
glass cylinder, or a jar with a wide mouth. A disc of cardboard, through which the delivery tube passes, rests on the top of the jar. Pour dilute hydrochloric acid down the funnel. During the effervescence a gas is given off and collects in the jar. When a burning taper is extinguished immediately it enters the jar, take out the delivery tube and put it into another jar. Cover the first jar of gas with a disc of card. In the same way collect several jars of the gas.

ii. **Properties of chalk gas.**—(*a*) Notice that the gas is (1) invisible, tasteless, and with but a faint smell ; (2) that it extinguishes a lighted taper ; (3) that it must be heavier than air, or it could not be collected in the way described.

(*b*) Pour the gas from one jar (*B*) into another (*A*), as shown in the diagram (Fig. 62), and test both jars by a lighted taper. It will be seen that the lower jar contains the gas.

iii. **Acid solution formed by chalk gas.**—Pour a little water made blue with litmus into a jar of the gas and shake it up. Some of the gas dissolves, and the colour of the solution turns a reddish purple. Boil the solution; the chalk gas is driven off and the blue colour is regained.

iv. **Magnesium burns in chalk gas.**—Place in jars of the chalk gas, lighted sulphur, a taper, carbon, and magnesium, and observe that all are extinguished with the exception of the magnesium, which continues to burn with a crackling sound. Examine the products of the combustion carefully, and note the white powder which formed the fumes during the burning, and also the numerous black specks (Fig. 63). Add a little dilute hydrochloric acid, notice that the white powder readily dissolves, but that the black specks remained undissolved.
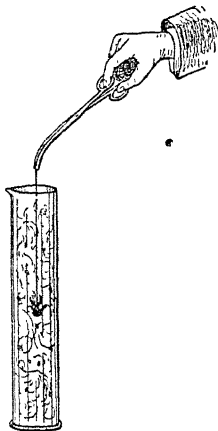
v. **Chalk gas obtained from marble.**—Collect in a similar manner a few jars of gas formed by the action of hydrochloric acid on marble, and, by any, or all, of the above tests, satisfy yourself that the gas evolved is identical with chalk gas.

FIG. 63.—Burning of magnesium in chalk gas.

Preparation of bottles of chalk gas.—The best way to prepare bottles, or jars, of the gas is to place pieces of chalk about the size of peas into a bottle fitted like that in Fig. 61. Dilute hydrochloric acid is poured down the thistle funnel, and when it comes into contact with the marble, or chalk, the gas is given off. Enough acid is poured in to cover the bottom of the funnel, so that the gas cannot escape up the funnel; it passes through the other tube in the cork. The gas given off is heavier than air, and can therefore be collected as shown in Fig. 61. As the gas accumulates in the jar, the air is pushed out at the top. After several bottles, or jars, have been filled, the properties of the gas can easily be examined. The gas may also, if desired, be collected over water as, although soluble,

its solubility is not sufficient to make this form of collection impossible, or inconvenient.

Properties of chalk gas.—An examination of the gas shows that it is colourless and has but a faint smell. As it is heavier than air it can be poured downwards just like a liquid (Fig. 62).

Chalk gas is slightly soluble in water, and the solution which is thus formed turns a blue litmus paper red, just as acids do. You will notice, however, that the red colour is not as bright as that produced by most acids, but has a port-wine colour ; the solution, although acid, is only a very weak acid. In naming the properties of this gas you must not forget that one which has been so frequently mentioned, namely, that it puts out the flame of a taper, or match ; it is consequently called a non-supporter of combustion. Although a non-supporter of the combustion of a match or taper—and indeed of most combustible bodies—it is very important to remember that magnesium will burn in chalk gas, doing so with a crackling noise, and with the production of both white fumes and black specks which mix with the white powder, and stick to the sides of the jar.

54. WEIGHT AND VOLUME OF CHALK GAS EVOLVED FROM ONE GRAM OF CHALK.

i. **Weight of chalk gas.**—Into a flask A (Fig. 64) fitted with india-rubber stopper and tubes, as shown, pour some dilute hydrochloric acid, and weigh the flask with the contained acid. Weigh a small tube B (i) empty, (ii) with some chalk. Place this in the flask, taking care not to allow the acid to reach the inside of the tube. Replace the cork
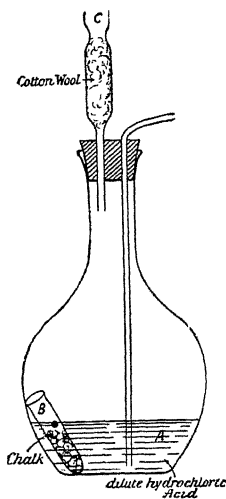
FIG. 64.—To illustrate Expt. 54, i.

and shake the flask so as to allow the acid and chalk to react.
The chalk gas formed escapes through the tube C in which
cotton-wool is packed to retain moisture. When the action is
at an end, draw a current of air through the flask in order
to displace the chalk gas, still left inside. Weigh the flask
again, and subtract the weight so obtained from that of the
flask + weight of tube and chalk. The resulting number is
evidently the weight of the chalk gas evolved. Calculate the
loss for 1 gram of chalk in the following manner .

Weight of tube and chalk, - -	=	6·364 gms.
Weight of tube, - - - -	=	4·177 gms.
Weight of chalk, - - -	=	2·187 gms.
Weight of flask and acid, - -	=	56·324 gms.
Weight of tube and chalk, - -	=	6·364 gms.
Total weight before action, -	=	62·688 gms.
Weight of flask, etc., after action, -	=	61·730 gms.
Loss = weight of chalk gas evolved	=	0·958 gm.

From 2·187 gms. of chalk, 0·958 gm. chalk gas is evolved.

$$,, \quad 1 \quad gm. \quad ,, \quad \frac{0·958}{2·187} \quad ,, \quad ,, \quad ,,$$
$$= 0·438 \ gm.$$

0·438 gm. of chalk gas are evolved from 1 gm. of chalk.

ii. **Volume of chalk gas.**—Determine the volume of chalk gas
evolved from a weighed quantity of chalk, in a manner similar
to that in which you determined the volume of hydrogen evolved
by the action of a known weight of zinc upon acid (Expt. 49, ii.).
The only difference is the substitution of chalk for the zinc,
whilst hydrochloric acid is preferable to sulphuric acid. One
precaution, however, should be taken. As chalk gas is some-
what soluble in water, some may dissolve in the water which
is being displaced, and consequently the volume obtained for
the gas evolved will be too small. In order to obviate this,
it is well to saturate beforehand the water in the displace-
ment jar with chalk gas so that no more can be dissolved
in it. This may be readily done by sprinkling in the water
a quantity of finely powdered chalk, and then adding some
hydrochloric acid. The chalk dissolves with evolution of
the chalk gas, which saturates the water. (The water may also
be covered by a layer of oil in which the chalk gas is not soluble,
but the above method is cleaner and quite as effective.)

Read the temperature and pressure, and as in Expt. 49, i. correct the volume found so as to obtain the volume under standard conditions, that is at 0° C. and at 760 mm. of mercury.

iii. **Repetition with marble.**—The above experiments may be repeated, employing marble in place of chalk. The same results are obtained, confirming the view that marble and chalk are identical in their chemical composition.

Density of chalk gas.—In a manner similar to that employed in the case of hydrogen, the weight and the volume of chalk gas, evolved from one gram of chalk when acted on by acids, may be readily determined. It will thus be found that from 1 gram of chalk about 0·44 gm. of chalk gas is given off. The value actually found will probably be a little less than this; but, if the chalk is perfectly pure, the above value will be obtained. The volume of chalk gas, when corrected so as to give the volume at standard temperature and pressure is 223 c.c. This volume is obviously the volume occupied by 0·44 gram, and we can hence readily calculate the weight of one cubic centimetre of the gas. The values actually obtained by experiment should be employed for the calculation—they will probably both be less than the above numbers owing to impurities in the chalk, but the value for the weight of 1 c.c. of the gas should be very nearly correct. Employing the above numbers it is evident that 223 c.c. of chalk gas weigh 0·44 gm. and hence that

1 c.c. of chalk gas weighs $\dfrac{0·44}{223}$ gm., that is 0·00198 gm.

The absolute density of chalk gas therefore is 0·00198 gm. per c.c. To determine the density compared with that of hydrogen it is necessary to divide this number by the weight of 1 c.c. of hydrogen, *i.e.* 0·00009 gm. We hence obtain the number 22, so that chalk gas is 22 times as heavy as hydrogen. It is easy to verify this by determining *directly* the density of the gas in the manner described in the next experiment

55. DIRECT DETERMINATION OF THE DENSITY OF CHALK GAS.

i. **Experimental determination.**—To a flask of about 200 c.c. capacity fit an india-rubber stopper with 2 holes. Through these pass two glass tubes bent at right angles ; one, A, passing to the bottom of the flask, the other, B, terminated just below the india-rubber stopper. Close the outer ends of the two tubes by means of pieces of india-rubber tubing fitted with pieces of glass rod (Fig. 65). Dry the flask thoroughly by heating it and aspirating through it a current of air, and, when the flask is perfectly dry. allow it to cool. Close the ends of the tubes A, B,

FIG. 65 — Flask for determination of gaseous densities.

and carefully weigh the flask and tubes. Next fill the gas with dry chalk gas, prepared by the action of hydrochloric acid on chalk or marble, and dried by passing through a tube of calcium chloride. The gas should enter the flask by the tube A, which reaches to the bottom of the flask (a gas lighter than air should be allowed to pass in by the short tube B. *Why?*) and should pass for several minutes so that the flask shall certainly be completely full of the chalk gas. The flask is then disconnected from the generating and drying apparatus, the ends of A and B closed, and the whole again weighed. The volume of the flask should be found by completely filling the flask and tubes with water, and pouring the water into a graduated cylinder. The temperature of the laboratory and the atmospheric pressure should be read, and the density calculated in following manner :

Temperature, 15° C. Pressure, 752 mm.

Volume of flask, -	-	-	-	-	224 c.c.
Corrected volume, -	-	-	-	210 c.c.
Weight of 1 c.c. of air,	-	-	-	0·00129 gm.[1]
Weight of 210 c.c. of air,	-	-	-	0·271 gm.
Weight of flask full of air, -	-	-	56·421 gms.
Weight of contained air,	-	-	-	0·271 gm.
Weight of flask if empty,	-	-	-	56·150 gms.

[1] If this weight has not been previously obtained during a Physics course, the flask may be first weighed full of dry hydrogen, the weight of each cubic centimetre of which is known to be 0·00009 gm.

Weight of flask full of chalk gas,	-		56·565 gms.
Weight of empty flask,	-	-	56·150 gms.

Weight of chalk gas, -	-	-	-	0·415 gm.

210 c.c. of chalk gas weigh,	-	-	0·415 gm.
1 c.c. ,, ,, weighs,	-	-	$\dfrac{0.415}{210}$ gm.

$$= 0·00198 \text{ gm.}$$

Density of chalk gas 0·00198 gm. per c.c.

Density of chalk gas determined directly.—The density of any gas may be determined directly in the manner explained in Expt. 55. It is necessary to determine the weight of a known volume of the gas. If a flask, the volume of which is known, could be weighed perfectly empty, and then full of the gas, the weight of the contained gas would of course be at once obtained. The flask employed, however, is full of air, and it is not easy to obtain it quite vacuous, so that the weight of the empty flask must be obtained by subtracting the weight of the contained air from the weight of the flask when filled with air. If the weight of one cubic centimetre of air has not been determined, the flask may be first filled with hydrogen and the weight of the contained hydrogen subtracted from the weight of flask and hydrogen. An additional weighing of the flask full of air will also suffice to give the weight of one cubic centimetre of air, so that in future, similar experiments, it will not be necessary to fill the flask with hydrogen. When the weight of the empty flask is known, it is evident that the difference between this weight and the weight of the flask full of chalk gas, must be equal to the weight of the chalk gas, and as the volume contained by the flask is known, the necessary connection between volume and weight has been obtained and the density of chalk gas can be estimated. It is thus found that chalk gas is 22 times as heavy as hydrogen. It is

hence nearly $1\frac{1}{2}$ times as heavy as oxygen, the density of which, compared with hydrogen, is 16, a result in accord with the previous indirect determination of the density.

QUESTIONS ON CHAP. XI.

1. A lump of lime and a piece of chalk are given you ; they are much alike in appearance ; how would you determine which was lime ?

2. Describe carefully what takes place when water is poured upon a piece of quicklime.

3. How would you prove that lime is soluble and chalk insoluble in pure water ?

4. When water is poured upon quicklime we believe that they unite together. What are the reasons for this belief ?

5. Knowing that chalk consists of lime and a gas which puts out flames, how would you set to work to find out the amount of the former in a lump of chalk given you ?

6. Describe the apparatus you would use for the preparation of chalk gas.

7. How would you collect some of the gas without a pneumatic trough ? Could hydrogen be collected in this way ?

8. If hydrogen is being collected without a trough, why do you not collect it in the same way as chalk gas ?

9. Describe experiments to show the high density of chalk gas and its effect on combustion.

10. Describe an experiment to determine the quantity of chalk gas which is evolved by the action of acid upon one gram of chalk.

11. Chalk is very commonly found in drinking waters. How do you account for this ?

12. What is the behaviour of (a) cold water, (b) hot water towards quick lime, common salt, saltpetre, and anhydrous copper sulphate respectively ?

CHAPTER XII.

EXAMINATION OF CHALK (*continued*).

THE nature of the change produced by heating chalk must be next examined, and the student should verify the statement that the product so formed is lime.

56. ACTION OF HEAT ON CHALK.

i. **Heating of chalk.**—Weigh a crucible and in it place some powdered chalk. Heat it in a muffle furnace at a bright red heat for a few hours, allow it to cool, and again weigh. Calculate the loss or gain for one gram of chalk.

ii. **Examination of product.**—Carefully examine the product and repeat all the experiments previously performed with lime (p. 134), carefully comparing the results.

iii. **Heating of marble.**—Repeat Expt. 56, i., ii., employing marble instead of chalk.

Action of heat on chalk.—It is found by heating a weighed quantity of chalk that the product formed, which examination proves is really lime, weighs considerably less than the original chalk, the loss being about 44 per cent., so that one gram of chalk loses 0·44 gram when strongly heated. It is clear that something has been driven out of the chalk by heating it. It was previously observed, that when lime is treated with acids, the products formed are identical with those obtained by the action of acids on chalk, but that no gas is evolved; whereas, with chalk a gas —chalk gas—is evolved with brisk effervescence. These

facts seem to indicate that the difference between lime and chalk might be that the chalk contains this gas united with lime, and that the heating of the chalk drives off this gas. As the gas is also driven off by the action of acids on chalk, if the above explanation is correct the same loss of weight should be found when chalk is heated as when it is acted on by an acid. The experiments performed have shown that this is actually the case, the loss being 0·44 gram per gram of chalk. It seems, therefore, that chalk is made up of lime and chalk gas, and that the latter may be driven off by the action of either acids or heat. It is hence possible that under suitable conditions lime and chalk gas might unite to form chalk ; both being somewhat soluble in water the addition of the two substances in solution would be most likely to effect this union.

57. ACTION OF CHALK GAS ON LIME-WATER.

i. Action of chalk gas on lime-water.—Prepare some chalk gas by the action of acid on chalk, and pass the gas from the delivery tube through some lime-water. Observe that a milkiness is produced, owing to the production of a white powder or *precipitate*, which disappears after a short time.
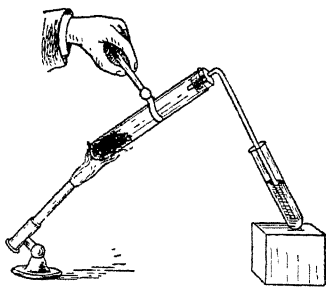
FIG. 66.—Chalk gas produced by heating chalk.

Boil the solution thus obtained, and notice that the milkiness again appears.

Filter the milky solution, and so obtain the white powder on a filter paper. Add a few drops of dilute hydrochloric acid to the powder. Notice the effervescence. Test the gas which is given off ; it puts out a flame.

ii. Action of gas from heated chalk on lime-water.—Place some chalk in a hard glass tube, fitted with a cork and delivery tube. Heat very strongly by means of a blowpipe flame (Fig. 66), and pass the gas evolved into a test-tube of lime-water, and proceed as in Expt. 57, i.

iii. **Action of chalk gas on milk of lime.**—Mix some powdered lime with water so as to obtain a white, milky liquid which consists of a saturated solution of lime with an excess of suspended lime. Pass through this *milk of lime*, contained in a test-tube, a brisk current of chalk gas for about half an hour. Filter the liquid and examine the product, comparing carefully the action of acids upon it and upon chalk.

Effect of chalk gas on lime-water.—When chalk gas is passed through a clear solution of lime (called lime-water) a white precipitate is obtained. This disappears when the gas passes through for a longer period, owing to the fact that the precipitate is soluble in a solution of chalk gas. On boiling the liquid, however, the chalk gas is driven off from the solution and the white solid is again precipitated. A larger quantity of the precipitate may be obtained if milk of lime is employed instead of lime water, and it may then be examined, when it is found to be chalk, effervescing with acids and giving the same products as were previously obtained from chalk. The gas given off from chalk by heating may be similarly treated and so compared with chalk gas, and it is found to behave in a precisely similar manner and to be identical with the gas driven off from chalk by the action of acids. If sufficient gas can be obtained by heating chalk, the gas may be collected over water and its properties in other particulars compared with those of the gas obtained by the action of acids (Expt. 53). The action on lime water is, however, the most certain test for this gas, and the production of the milkiness in lime water may be regarded as a sure evidence of the presence of the chalk gas. We may now write :

Chalk is composed of lime and chalk gas.

and also

In solution, lime and chalk gas unite to form chalk.

Some lime will also slowly unite with chalk gas even when not dissolved, and if exposed to chalk gas for a con-

siderable time will be found to contain a fair quantity of chalk.

58. COMPOSITION OF CHALK GAS.

i. **Burning of magnesium in chalk gas.**—Repeat Expt. 53, iv., and to the jar in which the magnesium has burnt add a little dilute hydrochloric acid ; shake well and pour into a test-tube. Observe that the white powder has dissolved and the black specks fall to the bottom of the tube. Pour off as much as possible of the liquid, leaving the black specks in the tube ; add a little water and again pour off. Heat and continue to heat after the liquid has boiled away ; observe that the black specks glow and ultimately disappear. Carefully consider what this experiment suggests regarding the composition of the chalk gas.

ii. **Burning of magnesium in oxide of carbon.**—Fill a jar with oxygen and burn in it a piece of carbon, employing a deflagrating spoon, which should be kept tightly pressed on the top of the jar. When the carbon ceases to glow brightly, remove the deflagrating spoon and close the mouth of the jar by a greased glass plate. Light a piece of magnesium ribbon and lower it when burning into the jar of oxide of carbon. Observe that it continues to burn. Compare carefully this experiment with the burning of magnesium in chalk gas, examining the products in a similar manner.

iii. **Action of oxide of carbon on lime-water.** – Prepare a jar of oxide of carbon as in Expt. 58, ii., and pour into the jar a little lime-water. Shake well, and observe the formation of the white precipitate. Pour the milky liquid into a second jar of oxide of carbon, and observe that it again becomes clear. (If necessary a third jar may be employed. The important point to be observed is that at first a milkiness is produced which, by the further action of oxide of carbon, disappears ; the two actions may of course take place in the first jar).

Boil the clear liquid and observe the reappearance of the precipitate, *i.e.* the chalk.

Add a little acid and observe that the precipitate dissolves with slight effervescence.

Chalk gas is oxide of carbon.[1]—When magnesium burns in chalk gas two products are obtained, (1) white fumes,

---

[1] In order to avoid misapprehension it is well to note that this gas which is here called ' oxide of carbon ' is also, and, more properly, called *carbon dioxide,* and also *carbonic acid gas.*

and (2) black specks. The fumes resemble those formed
by the combustion of magnesium in air or oxygen, and if
they are really of the same nature it is clear that the chalk
gas must contain oxygen   Sulphur and other combustible
substances, however, will not burn in the gas ; and, therefore,
if oxygen is present it must be already combined with some
other substance, from which it is only with difficulty
liberated.   The black specks may perhaps be this other
substance, and it is found that they are insoluble in water or
acid, and that they glow and disappear when strongly
heated.   In this they resemble carbon, so that it seems
probable that the chalk gas contains oxygen and carbon, or
that it is *oxide of carbon*.   In order to see whether this is
really the case, oxide of carbon should be compared with
chalk gas.   Oxide of carbon may be readily obtained by
burning carbon in a jar of oxygen.   Magnesium burns in
oxide of carbon with similar white fumes and black
specks as are obtained by its combustion in chalk gas.
Oxide of carbon has also the same action on lime water,
giving a precipitate of chalk, which is redissolved when excess
of the gas is passed in.   It may be further noticed that the
action of both gases on litmus has been previously found to be
identical (Expts. 38, vi., 53, iii.).   The gases are indeed
identical, and repetition of any tests would all lead to the
same result, viz. that *chalk gas* is *oxide of carbon*.   So that
we may now write :

Chalk consists of lime and oxide of carbon.

QUESTIONS ON CHAP. XII.

1. A sample of chalk yielded 132 grams of quicklime when strongly
heated ; what was the weight of the sample taken ?
2. Calculate the weight of lime and carbon dioxide respectively in
a cwt. of chalk.
3. Of what substances does chalk consist ?   Give full reason for your

answer.   Give the names of substances having a similar composition to chalk.

4. What happens when chalk is exposed to great heat? Describe a different method of obtaining one of the products.

5. When carbon dioxide is passed into lime water, the latter becomes first turbid and afterwards clear.   Explain the chemical nature of the causes of these changes.

6. How would you show that chalk consists of lime and carbon dioxide?

7. If you were given samples of limestone and rock salt, how would you proceed to obtain from them pure specimens of calcium carbonate and common salt respectively?

8. When oxygen is led over strongly heated charcoal and then into lime-water, a white precipitate is produced.   How would you set to work so as to place beyond doubt the nature of the precipitate?

CHAPTER XIII.

OXIDE OF CARBON IN THE ATMOSPHERE.

59. RECOGNITION OF ATMOSPHERIC OXIDE OF CARBON.

i. **Oxide of carbon is produced by a burning candle.**—(*a*) Burn a candle or taper in a clean, dry, white bottle (Fig. 40). After the flame has been extinguished, withdraw the taper. Pour a little freshly made lime-water into the bottle and shake it up. Notice the milkiness of the lime-water.

(*b*) Cut a long, thin chip of wood, hold it in the flame of a

FIG. 67.—Oxide of carbon produced by breathing.
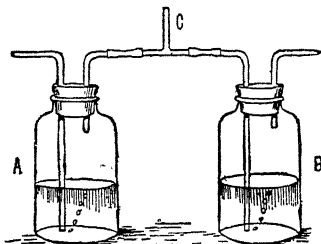
FIG. 68.—To illustrate Expt. 59, iv.

laboratory burner until it burns brightly, then thrust it into a cylinder or bottle, the bottom of which is covered with lime-water to the depth of about an inch. When the stick ceases to burn, withdraw it and shake the lime-water.

ii. **Oxide of carbon produced by breathing.**—(*a*) Blow through a piece of glass tube into some clear, freshly-made lime-water contained in a wine glass, or tumbler. Milkiness is at first produced, but if the blowing is continued long enough it by and by disappears (Fig. 67).

(*b*) Repeat the preceding experiment by blowing air from a bellows instead of from the lungs. Notice that this unbreathed air has not the same effects upon lime-water as breathed air, inasmuch as the lime-water is only very slowly turned milky.

iii. **The air contains oxide of carbon.**—Pour some clear lime-water into a blue dinner plate, or some other shallow vessel of a dark colour. Leave it exposed to the atmosphere for a little while. Notice the thin white scum formed on the top. The oxide of carbon in the air has turned the top layer of liquid milky.
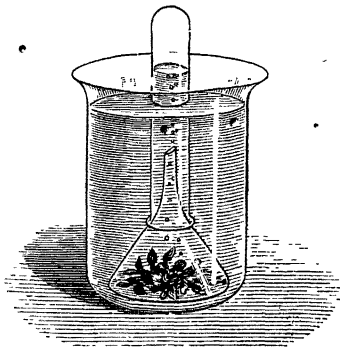
Fig. 69 —Production of oxygen by the action of plants.

iv. **Character of air changed by breathing.**—Fit two bottles with corks and tubes, as shown in Fig. 68. See that the corks are air-tight. Put some clear lime water into each bottle. Place the tube *C*, or an india-rubber tube leading from it, in your mouth. When you suck at the tube, air is drawn in through the glass tube which dips into the lime-water in the bottle *A*. When, however, you blow instead of sucking, your breath passes out through the tube which dips into the lime-water in the bottle *B*. Notice that the lime-water in *A* remains almost clear, but that in *B* is rendered milky by the air you breathe out. You thus see that fresh air has little effect upon lime-water, but breathed air quickly turns clear lime-water milky.

v. **Oxygen produced by action of plants.**—Take a bunch of fresh watercress, water crowfoot, or any aquatic plant, and put it into a beaker, or glass jar, completely filled with water saturated with oxide of carbon. Cover the plants with a funnel nearly as wide as the jar, as shown in Fig. 69. Fill a test-tube

with water, and invert it over the funnel.   If properly managed there should at first be no gas in the test-tube.   Place the jar in bright sunlight for an hour or two, and then examine it.   You will notice bubbles of a gas have collected at the top of the tube. Test the gas with a glowing splinter of wood.   It is found to be oxygen.[1]

vi. **Plants in sunlight and in darkness.**—Repeat the last experiment, but, instead of putting the bottle in bright sunlight, place it in the dark.   Observe that in such circumstances no bubbles of oxygen are formed.

Oxide of carbon formed by burning a candle.—The lime-water test for oxide of carbon is a very convenient one, and by its means we are enabled to detect with ease the presence of this gas.   We find that if a taper, a piece of wood, coal-gas, or oil, etc., burns in an enclosed space oxide of carbon

[1] A suitable form of apparatus for demonstrating to a class the fact that plants evolve oxygen, is the following (which appeared as an abstract in *The School World*) : A receiver holding two or three litres is employed,
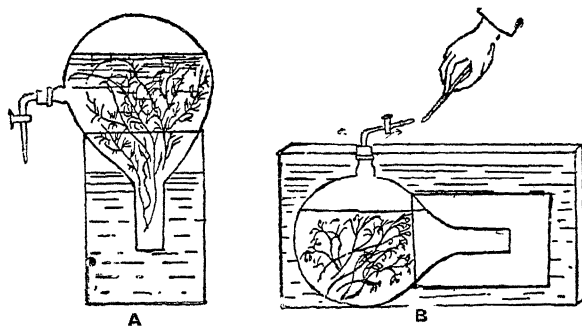
FIG. 70.

and into this a considerable quantity of an aquatic plant is introduced (Fig. 70 A).   The water is thoroughly charged with oxide of carbon, and the plants are then exposed to the sunlight.   Little streams of gas are seen to pass upward from various points, and when sufficient gas has collected at the top of the flask, the latter is immersed in a tank of water in a horizontal position in such a manner that the gas is directly under the opening (Fig. 70 B).   On turning the stop-cock and applying a splinter of wood with a spark on the end of it, the gas will be found to be oxygen.

is produced, for lime-water is soon turned milky if shaken up in the jar in which the burning takes place.

Oxide of carbon given off in breathing.—If a person blows with the mouth into clear lime-water, the lime-water is turned milky. This is another important fact. It is clear that oxide of carbon escapes from our lungs in breathing. And so it does from every animal. Not only, then, do all cases of ordinary burning cause the addition of oxide of carbon to the air, but also every act of breathing. It does not matter how small the animal is, all the time it is alive it is continually adding to the atmosphere a certain amount of this gas, the presence of which can be detected by its action in causing turbidity in clear lime-water.

Oxide of carbon in the atmosphere.—It is clear that inasmuch as all animals are continually pouring into the atmosphere a quantity of the oxide of carbon, their action being supplemented by coal fires, the burning of oil, gas, etc., the atmosphere must contain a quantity of this gas. That this is so is seen by Expts. 59, ii., iii. ; air drawn through lime-water slowly causes milkiness, whilst when exposed to the air in a shallow vessel (so as to expose a large surface) lime-water soon becomes coated with a film of chalk. The quantity of oxide of carbon in the air is, however, very small proportionally, about 0·03 per cent. or 3 parts in 10,000. This appears surprising when the enormous number of agencies by which it is produced is considered.

Action of plants.—One reason why there is never very much oxide of carbon in the air out of doors is because there are agencies continuously at work getting rid of this gas. The chief purifying agent is the green parts of plants which occur everywhere.

When fresh watercress is put into a bottle completely full of water containing oxide of carbon in solution, and the bottle is inverted in a basin of water without allowing air to get into the bottle, it is found that, when the bottle and its

contents are exposed to bright sunlight, bubbles of gas collect at the top of the bottle. These bubbles, when tested, are found to be pure oxygen. If, however, the bottle with the cress in it is kept in the dark, no bubbles or oxygen collect. Or, if a bottle of water in which oxide of carbon is dissolved be put in the sun, without any watercress, no oxygen collects in the top of the bottle.

In other words, two things are necessary for the formation of the bubbles of oxygen collected from the green plants as described. They are (1) the green vegetation, (2) the sunlight. The same conditions have been found always to hold true, thus proving that *green plants in the presence of bright sunlight have the power of turning oxygen out of oxide of carbon*. They keep the carbon for themselves, and it helps them to grow.

The balance of nature.—The wonderful relation between animals and plants in regard to their action upon the air, is a beautiful example of the way in which nature provides for one class of its creatures even out of the refuse of another. Animals by their breathing are continually using up the oxygen and returning to the atmosphere the compound oxide of carbon—a gas which is fatal to animal life. On the other hand, plants take in the oxide of carbon, and, in the presence of sunlight, the green material in their leaves can split up the gas into the two simpler substances of which it is built, namely, carbon and oxygen. The carbon the plants keep for themselves, the oxygen they do not use, and return to the atmosphere again. In this way plants help to keep the atmosphere in a suitable condition for the healthy life of animals.

It must not be thought, however, that this is the only action of plants and that they do not require any oxygen. They also, like animals, require oxygen, but their action in this respect is more than compensated in sunlight by their decomposition of the oxide of carbon.

QUESTIONS ON CHAP. XIII.

1. Give a brief account of the part played by carbon dioxide in the economy of nature.

2. By what means would you prove the presence of carbon dioxide in atmospheric air and expired air?

3. If two bottles were given you, in one of which phosphorus had been burnt, and in the other a candle, how could you decide which bottle was used for the phosphorus?

4. Describe a series of experiments you would make to illustrate the differences in the composition of ordinary air, dissolved air, and expired air.

CHAPTER XIV.

COMPOSITION OF OXIDE OF CARBON.

60. EXPERIMENTAL DETERMINATION.

i. **Passing oxygen over a weighed amount of carbon.**—Place a convenient amount of charcoal [1] in a wide, hard glass tube fitted with smaller tubes passing through india-rubber stoppers. It is convenient to place also loose plugs of asbestos between the charcoal and the stoppers. Weigh the tube with its stoppers and contents. Arrange three U-tubes as shown ; the first con-
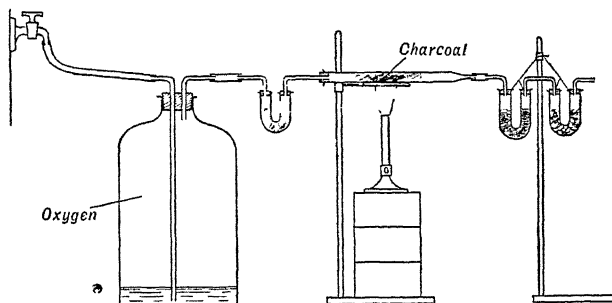
FIG. 71.—Apparatus for determining quantitatively the composition of oxide of carbon.

taining calcium chloride or sulphuric acid (to dry the gas) the second a strong solution of caustic soda, and the third lumps of the size of a pea of the same compound. Weigh the two last U-tubes with their contents and fittings. Join the apparatus as shown in Fig. 71. Heat the charcoal in the hard glass tube to redness, and pass dry oxygen over it from a gas-holder previously

[1] Charcoal employed in quantitative work must always be strongly heated beforehand, as it almost always contains a considerable quantity of moisture.

filled with oxygen. After the experiment has gone on for ten minutes, disconnect from the oxygen supply, and remove the burners. Disconnect the apparatus, and again weigh the U-tubes. When the hard glass tube is cool, weigh it again also. Notice that the U-tubes are heavier, whilst the hard glass tube and its contents weigh less than before. Calculate the quantity of oxide of carbon produced from one gram of charcoal For every gram of charcoal used up, it will be found that 3⅔ grams of oxide of carbon are absorbed by the caustic soda.

ii. **Composition of oxide of carbon, second method.**—(a) Weigh a small crucible, place in it a few pieces of charcoal which have been previously thoroughly heated to drive off any moisture they contain. Obtain thus the weight of the charcoal. Heat the crucible strongly until the charcoal has completely burnt away, leaving only a white ash in the crucible. Weigh the crucible and ash, and by subtracting the weight of the crucible find the weight of the ash. Calculate the percentage of ash in the charcoal.

(b) Weigh a hard glass tube containing a quantity of powdered oxide of lead, *i.e.* litharge, and add to it some of the charcoal employed above, but in the powdered form, (the weight should only be about one hundredth of that of the oxide of lead). Weigh the tube again and so obtain the weight of the charcoal. Thoroughly mix the two powders in the tube, taking care, however, to lose none, and then heat strongly for some time. The tube will be found to weigh less than before. The loss is the oxide of carbon, formed from the carbon and the oxygen of the oxide of lead. Calculate the oxide of carbon formed from one gram of carbon, taking care to allow for the ash, thus :

Weight of tube, lead oxide, and charcoal,   79·895 gms.
Weight of tube and lead oxide,   -   -   79·784 gms.

Weight of charcoal   -   -   -   0·111 gm.
Percentage of ash previously found, -   =   8·3 per cent.
Pure carbon,   -   -   -   -   -   = 0·111 × ·917 gm.
                                          = 0·102   gm.
Weight of tube, lead oxide, and charcoal,   -   = 79·895 gms.
  „   „   „   „   after heating, = 79·522 gms.

Weight of oxide of carbon evolved,   -   0·373 gm.

0·102 gm. carbon forms 0·373 gm. oxide of carbon.

1 gm. carbon   „ $\dfrac{0·373}{0·102}$ gms.   „   „

= 3·66 gms.

Composition of oxide of carbon.—We may now consider how the quantity of oxygen and carbon present in oxide of carbon can be determined. To do this we must find the weights of two of these three substances, (?) carbon, (2) oxygen, (3) oxide of carbon. The weights most easy to obtain are those of the carbon and the oxide of carbon ; the weighing of the carbon offers no difficulties, but to find the weight of the oxide of carbon it will be necessary to absorb it by some suitable substance. The most convenient is the substance known as caustic soda, a strong solution of which rapidly absorbs the gas. If, therefore, oxygen be dried and passed, in a gentle current, over charcoal strongly heated in a hard glass tube, oxide of carbon will be produced, and can be absorbed in U-tubes containing caustic soda. It will be found that the tube of carbon loses in weight, the loss being the quantity of carbon which has been burnt; the U-tubes, however, gain in weight, the gain being the quantity of oxide of carbon produced. The following in an example of such an experiment.

Weight of tube of carbon before experiment, $= 35\cdot672$ gms.
Weight of tube of carbon after experiment, $= 35\cdot438$ gms.

Loss = carbon burnt,   -   -   - $= 0\cdot234$ gm.

Weight of U-tubes after experiment,   - $= 42\cdot970$ gms.
Weight of U-tubes before experiment,   - $= 42\cdot114$ gms.

Gain = oxide of carbon,   -   -   - $= 0\cdot856$ gm.

$0\cdot234$ gm. of carbon forms $0\cdot856$ gm. oxide of carbon.

$$\text{I gm.} \quad ,, \qquad ,, \qquad \frac{0\cdot856}{0\cdot234} \quad ,, \qquad ,, \qquad ,,$$
$$= 3\cdot66 \text{ gms}$$

The determination may also be made by heating a known weight of carbon with some oxide which can give up oxygen to the carbon—lead oxide will serve the purpose. If this is done, then the total loss of weight is the weight of the oxide of carbon which has been evolved.

Care must be taken, however, to allow for the ash which is always contained in ordinary pieces of charcoal, and the quantity of ash must be found by means of a preliminary experiment. This having been done the results are obtained in the manner shown in Expt. 60, ii.

It is hence found that 1 gram of carbon forms 3·66 grams of oxide of carbon. We have already found that 1 gram of chalk contains 0·44 gram of oxide of carbon ; this, therefore, is made up of $\dfrac{0·44}{3·66}$ gm. of carbon and the remainder oxygen. This will be found to yield 0·12 gram carbon and 0·32 gram oxygen, so that we now know 1 gram of chalk is composed of 0·56 gram of lime, 0·12 gram of carbon, and 0·32 gram of oxygen.

Oxide of carbon contains its own volume of oxygen.—It has been found that 1 litre of oxide of carbon weighs 1·98 grams (p 141), and it is now known that of this $\dfrac{1}{3·66}$ is carbon and $\dfrac{2·66}{3·66}$ is oxygen. The weight of oxygen in one litre of oxide of carbon is therefore $\dfrac{1·98}{1} \times \dfrac{2.66}{3·66}$ gms. that is 1·44 gms. But this has been previously found to be the weight of 1 litre of oxygen. It hence appears that one litre of oxide of carbon contains one litre of oxygen, in other words, oxide of carbon contains its own volume of carbon. There is, therefore, no change in volume when carbon burns in oxygen, the oxide of carbon produced occupying the same volume as the oxygen from which it was produced.

(This last statement may be readily verified experimentally, and it affords a good practical exercise for the student to devise apparatus and carry out the experiment.)

Elements and compounds.—It will probably have been noticed at this stage that a number of substances have been examined which have been altered or changed into new

substances by the action of acids or of heat. Thus, by
heating we have obtained lead oxide from lead, copper
oxide from copper, potassium chloride and oxygen from
potassium chlorate, oxide of carbon from carbon, lime and
oxide of carbon from chalk, etc. Now if these changes are
examined it will be found that they may be divided into two
classes.

First, those in which a new product is formed which
weighs more than the original, so that the original substance
has combined with some other substance to produce the
new compound—such a process is a **combination.**

In the second class of reaction the substance has yielded
two or more products each weighing less than the original,
the process has been one of **decomposition.**

Now some of the substances examined have never been
found to undergo the second class of change. Thus, from
copper we cannot get two or more products each weighing
less than the original copper, and together making up the
weight of the copper, as we did in the case of chalk.
Similarly in the case of zinc—the zinc oxide obtained by
heating zinc in air, the zinc sulphate obtained by the action
of sulphuric acid on the metal, etc , all weigh more than the
original zinc.   Such substances which cannot be decom-
posed into two, or more simpler substances are termed
elements.   Thus, zinc, lead, copper, tin, mercury, oxygen,
hydrogen, carbon, iron, nitrogen are all regarded as elements.
Should at any time some chemist succeed in decomposing
one of these elements into two or more substances, then that
so called element would no longer be regarded as such. Up
to the present, however, none of those substances we regard
as elements has been so decomposed.

Substance which can be decomposed are called **compounds,**
and examples which we have studied are oxide of carbon,
oxide of copper, and in fact all the substances known as
oxides, potassium chlorate, chalk, etc.

Such compounds may be divided into classes—thus all those which consist of oxygen united with another element are termed *oxides*, and all substances which resemble chalk in containing oxide of carbon united with another substance, the analogue of the lime, and give off this oxide of carbon when treated with an acid are termed *carbonates*. Some of these carbonates should next be examined.

QUESTIONS ON CHAP. XIV.

1. What is the weight of oxygen contained in 20 litres of oxide of carbon?

2. Calculate the volume and weight of oxide of carbon produced by the combustion of 5 grams of carbon.

3. What (1) weight (2) volume of oxygen is necessary for the complete combustion of 15 grams of carbon?

4. Calculate the weight of carbon present in 1 cwt. of chalk.

5. Explain clearly the meaning of the terms, Element and Compound. Why would you consider chalk to be a compound, but zinc an element?

What reasons have you for stating that oxide of carbon contains its own volume of oxygen? Devise a method for testing this statement experimentally.

7. Assuming that the carbon used in question 3, contained 9·4 per cent. of ash, what volume of oxygen would then be necessary?

8. Calculate the percentage composition of oxide of carbon.

CHAPTER XV.

CARBONATES.

61. EXAMINATION OF TYPICAL CARBONATES.

i. **Sodium carbonate, soda crystals, or washing soda.**—(*a*) Examine carefully according to the scheme on p. 13, some washing soda crystals. Observe that when heated the crystals give off water of crystallisation and form a white powder. Try the effect of acids on the crystals and on this white powder. Observe the effervescence, and test by means of lime-water the gas evolved. Observe that this gas is oxide of carbon, and the substance is therefore a carbonate.

(*b*) Heat the white powder strongly and again test with acid. Observe that the gas is still evolved and, therefore, has not been driven off by the heating.

ii. **Copper carbonate.**—(*a*) Examine according to the scheme on p. 13. Note very carefully the action of heat, and test with lime-water the gas evolved. Observe that the same gas is evolved when the substance is treated with an acid.

(*b*) Examine carefully the black powder left after heating. It dissolves in acid *without* effervescence, the oxide of carbon has therefore been driven off by heat.

(*c*) Compare it with other black powders you know of, as for example soot, oxide of copper, manganese dioxide, observing carefully the action of acids on each. Satisfy yourself that the black powder is oxide of copper.

iii. **Lead carbonate.**—Examine lead carbonate in precisely the same manner as the copper carbonate. Compare carefully with oxide of lead (litharge) the product left on heating. The effect of strongly heating is so characteristic that no doubt as to the identity of the two will remain.

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iv. **Zinc carbonate.**—(*a*) Examine as before. Note very carefully that the substance left after heating is yellow when hot, but becomes white again on cooling. This was found (p. 31) to be characteristic of oxide of zinc.

(*b*) Compare this product with oxide of zinc, noticing the ac oistened litmus paper,—it is distinctly alkaline. Compare the action of sulphuric and hydrochloric acid on both.

Composition of other carbonates.—All carbonates do not decompose on heating. Thus, the ordinary soda crystals, or washing soda, may be readily proved to be a carbonate as they give off oxide of carbon copiously when treated with an acid. These crystals contain water of crystallisation which is given off when the crystals are heated, and a white powder remains. It might be thought that this would be the analogue of lime, but treatment with an acid shows that this is not the case, and that the white powder still contains the oxide of carbon, the only thing driven off by heat from the crystals being the water. The white powder left is the dry, or *anhydrous*, sodium carbonate.

Copper carbonate is a green powder which is insoluble in water. In acids, however, it dissolves very readily with a copious evolution of oxide of carbon and the production of blue or green liquids. When heated, copper carbonate loses oxide of carbon, doing so far more readily than chalk in similar circumstances, and leaves behind a black powder. This black powder may be readily proved to be oxide of copper by comparison with some of the latter compound. Copper carbonate, therefore, consists of oxide of carbon and oxide of copper.

Lead carbonate is a white insoluble powder which, like copper carbonate, gives off oxide of carbon when treated with an acid. When acted upon by either sulphuric or hydrochloric acid, however, the other product formed is almost insoluble in cold water, and the lead carbonate appears therefore to remain undissolved. On heating, lead

carbonate also readily decomposes, with the evolution of oxide of carbon, while litharge, or oxide of lead, remains. This litharge is readily recognised by fusing, when heated, to a dark brown liquid, which on cooling changes to a bright yellow, glassy solid. Lead carbonate, therefore consists of oxide of carbon and oxide of lead.

Zinc carbonate is also a white, insoluble solid which readily dissolves in acids with the evolution of oxide of carbon. On heating, it loses oxide of carbon and leaves a solid which is readily recognised as oxide of zinc. This, like lime, is alkaline and turns moistened red litmus paper a blue colour. Zinc carbonate is, therefore, composed of oxide of carbon and oxide of zinc.

Lime is a metallic oxide.—A number of carbonates have now been examined; in their general behaviour they resemble chalk and one another. They all give off oxide of carbon when treated with an acid, and those which are decomposed by heat also give off the oxide of carbon when heated. Where the product left on heating can be recognised, it is found in each case to be a metallic oxide Thus, copper oxide, lead oxide, and zinc oxide have been obtained by heating the carbonates examined. Now, as chalk in similar circumstances loses oxide of carbon and leaves lime, we are forced to the conclusion that lime also is a metallic oxide. This is supported by the fact that lime is alkaline, as the other alkaline compounds examined—for example, oxide of sodium, oxide of magnesium, and oxide of zinc—are also metallic oxides. The metal contained in lime is termed **calcium**, and although lime is so plentiful, yet it is so difficult to decompose it, that the metal itself is very expensive. We can now write:

Chalk consists of oxide of carbon and the oxide of a metal.

The metal has been named calcium, and chalk may hence be termed **calcium carbonate.**

QUESTIONS ON CHAP. XV.

1. What is the chemical nature of chalk and lime? How could you obtain lime from chalk and *vice versa*?

2. What are carbonates? In what respects do all carbonates resemble chalk and one another? In what respects do they differ?

3. What reasons have you for believing that lime is an oxide of a metal?

4. Describe the effect of heat on copper carbonate and on lead carbonate.

CHAPTER XVI.

COAL GAS, AND THE BUNSEN BURNER.

THE student will now be better able to understand the Bunsen burner which he has been using so frequently during his studies. To do so it is necessary that the gas itself, coal gas, should be studied also.

62. COAL GAS.

i. **Distillation of coal.**—Heat some coal dust in a hard glass tube to which an india-rubber stopper and delivery tube are attached. After a time apply a lighted taper to the end of the delivery tube, and satisfy yourself that an inflammable gas is given off.

ii. **Coal gas is lighter than air.**— Show, as in Expts. 45, ii., iii., that coal gas escapes from a cylinder held mouth upwards. It is lighter than air.

iii. **Coal gas is insoluble in water.** —Invert a cylinder of coal gas over water, and observe that the water does not rise. Closing the end of the jar tightly by the hand, shake well with a little water and again place over water. Observe that the water does not rise in the jar; the gas is therefore insoluble.

FIG. 72.—Apparatus for combustion of coal gas in a cylinder.

iv. **Oxide of carbon produced by the combustion of coal gas.**—By means of a bent glass tube and a piece of india-rubber tubing attached to the gas supply, as shown in Fig. 72, allow a small

jet to burn in a cylinder. Notice the formation of drops of liquid on the sides of the cylinder. Remove the gas jet Add a little clear lime-water, and shake it up and down the cylinder. The lime-water is turned milky, showing the presence of oxide of carbon.

v. **Water produced by combustion of coal gas.**—Allow a gas jet to burn for some time under a retort kept cool by a stream of water exactly as in Expt. 46, 1. Collect the liquid produced, and as in Expt. 46, ii., prove that this liquid, which may require filtering, is water.
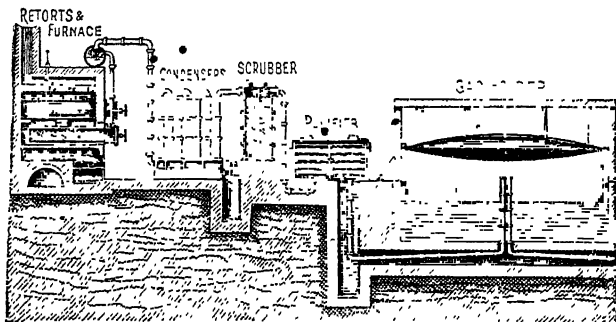
FIG. 73 —Manufacture of coal gas.

Coal gas.—Coal gas is produced by distilling coal in large iron retorts. Other products, such as tar, also result, but these are separated from the gas, which, after purification from certain undesirable constituents, is passed into large receivers, the *gasometers*, from which it can be used as required (Fig. 73) It is an invisible gas with a well-known odour. It is considerably lighter than air, but still much heavier than hydrogen. When it burns, oxide of carbon is produced, which may be detected by its action on lime water, while water is also formed. It is evident, therefore, that the coal gas must contain *carbon* and *hydrogen*. These are indeed, its chief constituents as it consists chiefly of compounds termed *hydrocarbons* which contain only carbon and hydrogen.

63. FLAME AND ILLUMINATION.

i. **A candle flame.**—Examine the flame of a candle ; notice that it is divisible into three well-marked zones. Outside is a dark yellow layer ; in the middle a highly luminous layer can be made out ; the interior of the flame consists of a bluish part which gives very little light.

ii. **The nature of the three zones.**—(*a*) Take a short piece of glass tubing, about the thickness of a pencil, and open at both ends, arrange it so that one end is in the inner blue zone, while the tube itself slants away from the flame (Fig. 74). Apply a lighted match to the end of the tube away from the candle ; observe that a gas which catches fire is issuing from the tube. In other words. the inner zone of a candle flame is made up of unburnt combustible gases.

(*b*) Lower a piece of cold, thick, white notepaper on to the flame ; after it has been there a second or two, raise it and examine the ring of soot which has been deposited. Round the ring of lamp-black is a second ring, where the paper is slightly charred. Repeat the experiment, bringing the piece of paper into the flame from the side, about half-way up the inner bluish zone ; the

FIG. 74.—To illustrate Expt. 63, ii.

results thus obtained are perhaps more satisfactory. These results are explained by the presence of incandescent particles of carbon in the highly luminous zone. The ring of charring is caused by the outside layer, where the combustion of the vapours from the candle is complete. Place also a piece of glass tubing in the flame and note the deposit of soot on it.

iii. **A gas flame.**—Examine a gas flame. Three zones similar to those in the candle flame can be made out. Using a piece of glass tubing as before, lead some of the unburnt gases in the centre of the flame outside, and ignite them at the end of the tube. Prove the presence of particles of carbon in the middle zone by lowering a cold white plate upon the flame. Examine the deposit of soot.

iv. **Result of lowering the temperature of flames.**—(*a*) Make a short coil of stout copper wire, $\frac{1}{4}$-inch internal diameter. Pass it over the wick of a lighted candle without touching the wick. The candle is extinguished. This is due to the cooling effect of the wire, which conducts away the heat.

(*b*) Turn on but do not light, a gas-jet. Hold over it a piece of fine wire gauze and apply a light above the gauze. Notice that the flame does not strike through. Vary the experiment by lowering a piece of cold wire gauze upon an ordinary Bunsen flame.

v. **The Bunsen burner.**—(*a*) Examine a Bunsen burner, or, as it is often called, a laboratory burner. Compare its parts with Fig. 75. The metal tube, T, screws out, and reveals a small jet, J, through which the coal gas is supplied. One or two holes, A, can be closed by rotating the movable cylinder which fits round a slightly smaller fixed tube. Screw in the tube, T, open the holes, A, and connect the burner with the gas supply by means of a suitable length of india-rubber tubing. Turn on and light the gas. An almost colourless flame is obtained. It is a mixture of coal-gas and air which is burning. Close the holes, A, the flame immediately becomes luminous.
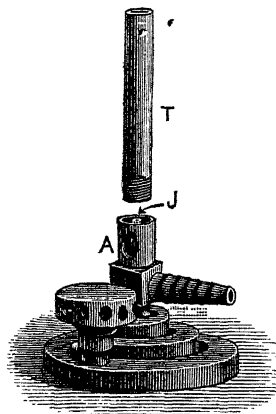
Fig. 75.—A Bunsen burner.

(*b*) Again obtain a non-luminous flame; lower a cold white evaporating basin into it, and observe there is no deposit of carbon. Close the holes, A, and lower the cold basin into the luminous flame you have now got; there is an abundant supply of soot.

(*c*) Blow some dust through the non-luminous flame of a Bunsen flame, and observe that it at once becomes much brighter. This can be readily done by placing some dust or any fine powder in a piece of glass tubing and then blowing it into the flame. There is a connection between the incandescent particles in a flame and its brightness.

The candle flame.—To obtain a candle flame, the wax of the candle has first to be melted ; it then soaks the wick and is converted into vapours which, like coal gas,

consist of carbon and hydrogen. These vapours burn to produce the flame.

The flame thus produced is not as simple as is first imagined. It consists of three zones, each having its own particular characters. In the centre of the flame is a bluish cone of unburnt vapours, just as they were obtained by the vaporisation of the liquid wax. Outside this is the layer of the flame to which all its brightness is due ; it is impregnated by particles of carbon, which the heat of combustion has made incandescent. This can be easily proved by lowering on to the flame any cold white surface, when the particles of soot are at once deposited. In this part of the flame there is not sufficient oxygen to combine with all the carbon, so that some remains unburnt. The part of the flame which adjoins the air is known as the zone of complete combustion ; this layer is not highly luminous, but it has the highest temperature of any part of the flame. Here the whole of the unburnt carbon unites with oxygen and is converted into oxide of carbon.

It must now be explained what, from the point of view of the chemist, a flame really is. When two substances unite to form a new compound, heat is generally produced, and this heat may be sufficient to raise the temperature of the substances and of the products of their combustion to such a degree that they become luminous. If, as in the case of the candle, gases are present among the products of combustion, they become luminous and pass away as flames. It may, in consequence, be stated that flames are gases produced by combustion, which have had *their temperatures raised so much that they have become self-luminous.*

If by any means the temperature is lowered, the flame is extinguished. The truth of this statement is at once demonstrated by lowering a cold coil of copper wire on to a candle flame without touching the wick, when the flame is put out. Or, the same thing can be demonstrated by lowering a piece of

cold wire gauze on to a non-luminous flame, such as that of a laboratory burner. In this case no flame appears above the gauze, because the heat is conducted away by the wire, and the temperature is thus lowered to such a degree that flame is impossible.

A gas flame.—The only important difference between a gas flame and that of a candle is that in the former the combustible compounds are already in the state of vapours, and have not to be converted into vapour by the heat of the flame. The same zones, however, may be made out in the flame.

Bunsen burner.—In the Bunsen burner there is a simple arrangement by which air is allowed to mix with the gas before the combustion. The result of this is that there is sufficient oxygen for the whole of the carbon of the components of the gas to be converted into oxide of carbon, and the luminous portion of the flame, which consisted of unburnt carbon, in consequence disappears. The flame becomes non-luminous, but, owing to the more complete combustion the temperature of the flame is much higher than that of the gas flame which is used for purposes of illumination.

QUESTIONS ON CHAP. XVI.

1. What does coal-gas consist of? What products are formed by its combustion, and how would you endeavour to prove your answer experimentally?

2. Write a little essay on "the burning of a candle."

3. Describe a Bunsen burner and the precautions that should be taken in using it.

4. Describe an experiment showing that water and carbon dioxide are formed in the combustion of a candle or of a gas jet. How can ordinary coal-gas be made to burn with a non-luminous flame?

5. Give the names of four common combustible substances. Explain as fully as possible what takes place when a gas jet is lighted.

CHAPTER XVII.

THE EXAMINATION OF SALT AND NITRE.

64. SALT AND SALT GAS.

i. **Action of strong sulphuric acid on salt.**—(*a*) Place a little salt in a test-tube, and add to it some strong sulphuric acid; warm slightly. Observe that a gas is evolved which possesses a powerful, pungent odour, and forms white fumes in the air, although it appears colourless in the tube itself.
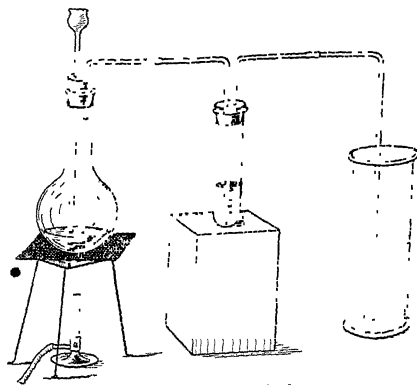
FIG. 76.—Preparation of salt gas.

(*b*) Plunge a lighted taper in the gas and notice that the flame is extinguished. Place a piece of moist, blue litmus paper at the mouth of the tube, and observe that it is turned red.

ii. **Preparation of salt gas.**—Fit up the apparatus shown in Fig. 76. Remove the india-rubber stopper of the flask and place in it a small quantity of rock salt in small pieces, or some

thoroughly dried common salt. Pour some strong sulphuric acid into the wash-bottle shown in the middle of the illustration. Re-insert the india-rubber stopper into the flask and pour down the acid funnel enough of a mixture[1] of strong sulphuric acid and water to cover the salt in the flask. Gently warm the flask. Collect jars of the gas which is evolved (after it has bubbled through the strong sulphuric acid in the wash-bottle and so become freed from water vapour) in gas jars, by downward displacement in the way the illustration makes clear. When each gas cylinder is full (this is probably the case when a blue litmus paper held just below the top of the outside of the cylinder is turned red), cover it with a ground glass plate, with the ground side underneath. As the cylinders are filled set them on one side for examination as presently described. Collect four jars of gas in this way.

iii. **Properties of salt gas.**— (a) Raise the glass plate from the first jar and plunge a lighted taper into the gas. The flame is extinguished and the gas does not burn. Quickly replace the glass plate.

(b) Into the same jar drop a piece of moistened, blue litmus paper and replace the glass plate. The paper is turned red showing the gas has acid properties. Notice carefully that the paper is not bleached.

(c) Observe the fumes which the gas forms with the air when the glass plate is removed from a cylinder full of the gas. This is due to the very strong power of absorbing moisture possessed by the 'salt gas.'

(d) Firmly pressing the glass plate, invert a cylinder full of the gas and place it upside down in a basin of water. Remove the plate when it is under water, and notice that water rushes up and completely fills the jar. If it does not completely fill the jar, it shows that the air in the jar was not altogether displaced by the gas when you should have filled it.

iv. **Solution of salt gas in water.**—Modify the apparatus shown in Fig. 76. Remove the gas cylinder and the delivery tube which dips into it. Pour out the sulphuric acid from the wash-bottle, thoroughly wash the bottle and half fill it with water. If there is still enough salt and sulphuric acid in the flask, again warm it gently and allow the evolved gas to bubble into the water. Notice that it is completely dissolved.

v. **Salt gas solution is hydrochloric acid.**—(a) Pour a little

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[1] One part of acid and one of water are convenient proportions. Be careful gradually to pour the acid into the water and not the water into the acid when mixing them, keeping the mixture well stirred throughout the process.

caustic soda solution into an evaporating basin, and add to it some of the salt gas solution until the liquid becomes acid. Evaporate the solution to dryness and taste the white solid which results, observing by the taste that it is ordinary salt.

(*b*) Repeat the experiment using hydrochloric acid instead of the salt gas solution, and note that salt again results.
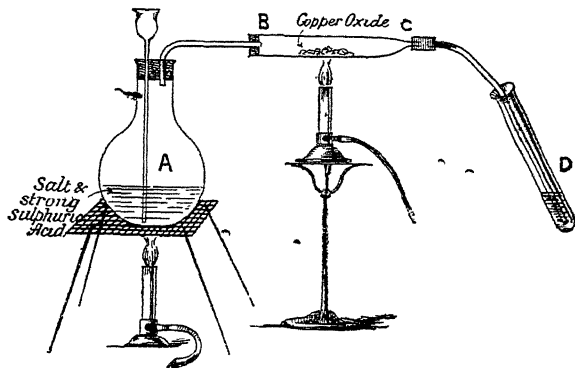
FIG 77.—To illustrate Expt. 64, vi. (a).

vi. (*a*) **Hydrochloric acid gas contains hydrogen.**—In the flask A (Fig. 77) place some salt and strong sulphuric acid   Pass the salt gas, or hydrochloric acid gas, as we can now call it, so obtained over heated copper oxide in the hard glass tube BC.   Observe that **water** collects in the test-tube D, and that the copper oxide is converted into a green substance.   Examine this green substance and see that it also gives off salt gas when acted upon by sulphuric acid.   As water is formed the salt gas must evidently contain hydrogen.

(*b*) Collect a tube of hydrochloric acid gas over mercury, and quickly introduce into it a piece of clean sodium.   Allow it to stand, and observe that the volume of the gas becomes ultimately reduced to one-half the original volume (correction being made for the difference in pressure), while further, the sodium gets covered with a white powder which you may satisfy yourself is salt.   Test the gas left with a lighted taper and see that it has the properties of hydrogen.

(*c*) Repeat Expt. 64, vi., (*a*), but employing zinc instead of the oxide of copper, and allowing the delivery tube to dip into a trough of water beneath an inverted test-tube of water. Observe that hydrogen collects in the test-tube.

Salt gas or hydrochloric acid gas.—When salt is acted on by strong sulphuric acid, a gas is evolved which possesses a strong, pungent odour, has acid properties and, though colourless itself, forms white fumes when it enters the air. The fumes are more evident in moist air, as is seen by breathing over the mouth of the test-tube from which they are being evolved. The gas readily dissolves in water, and the solution constitutes the hydrochloric acid of commerce, as may be readily proved by performing the same tests with the solution of salt gas and with the ordinary hydrochloric acid of the laboratory. As an example of one such test it may be stated that both liquids when treated with caustic soda form a solution of common salt   The solution of salt gas was, because of its preparation from salt, originally known as "Spirits of Salt." Another common name given to it is Muriatic Acid. The salt gas itself is called hydrochloric acid gas. It is colourless, will not allow things to burn in it, nor will it burn itself. As is seen by its action on blue litmus paper, it is strongly acid. It is heavier than air, and can consequently be collected by downward displacement.

Composition of hydrochloric acid.—You have already learnt that when the element copper is heated strongly in air it combines with the oxygen of the air to form a black compound called copper oxide, which compound is evidently entirely composed of copper and oxygen.

Now, if hydrochloric acid gas is passed over heated copper oxide, which can be conveniently arranged in a tube, the black oxide is changed into a green substance, and water is formed at the same time. Moreover, if the green substance be acted upon with strong sulphuric acid, hydrochloric acid gas is again formed, just as when salt is similarly treated.

As water is formed, the hydrochloric acid gas passed over the heated copper oxide must evidently contain the hydro-

gen which is necessary for the formation of the water, for copper oxide contains none.   In fact, the simplest explanation of the experiment is that the hydrogen of the hydrochloric acid gas combines with the oxygen of the copper oxide to form water, while the copper combines with the other constituent of the hydrochloric acid to form a green substance known as chloride of copper.

When metallic sodium acts upon hydrochloric acid gas contained in a tube over mercury, it is found after a short time that the volume of the gas is reduced to one-half and that the gas left in the tube is pure hydrogen.   At the same time, the sodium combines with the other constituent of the hydrochloric acid gas to form a white solid, which proves on examination to be common salt.   Zinc also, when heated, abstracts something from hydrochloric acid gas and hydrogen is liberated.

We have thus found that from hydrochloric acid gas, sodium abstracts something and leaves one-half the original volume of hydrogen.   Hence, also, salt consists of sodium with *something*, and this same material with hydrogen forms hydrochloric acid gas.   The name given to this material is **chlorine**, and we shall now endeavour to obtain this element from the hydrochloric acid gas.

65. PREPARATION AND PROPERTIES OF CHLORINE.

i. **Preparation of chlorine.** — In a fairly large flask (Fig. 78) put some manganese dioxide (a black powder, which also occurs naturally in compact masses known as *pyrolusite*).   Through the thistle funnel

FIG. 78.—Preparation of chlorine.

pour in sufficient hydrochloric acid (*i.e.* the strong acid used in the laboratory) to cover the oxide and see that the end of
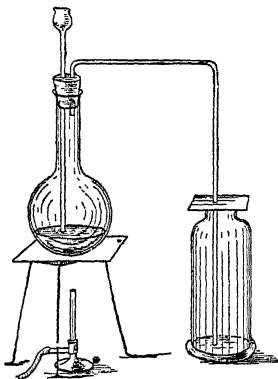
the thistle funnel is below the level of the acid. Observe that a dark, greenish-brown liquid results. Warm and observe the formation of a heavy, green gas which may be collected in the same manner as hydrochloric acid gas.[1] In this manner fill five jars, and close each jar tightly with dry ground glass-plates.

ii. **Bleaching power of chlorine.**—In the first jar place some moist, coloured rags, or flowers, and observe the bleaching. Writing in ordinary ink may also be bleached, but printers' ink cannot be bleached.

iii. **A taper burns in chlorine.**—Plunge a lighted taper in the second jar, and observe that the taper burns but with the formation of clouds of soot, and observe that water shaken up in the jar becomes acid.

iv. **Turpentine burns in chlorine.**—Moisten a piece of filter paper with turpentine, fold it and place it in a jar of chlorine. Observe the blackening and the final burst into flame, with the production of clouds of soot and also white fumes. Shake up the fumes with water and observe that the solution is acid.

v. **Burning of phosphorus in chlorine.**—Place in the jar, by means of a deflagrating spoon, some burning phosphorus, and observe it continues to burn with the formation of white fumes.

vi. **Burning of sodium in chlorine.**—Place in the gas some burning sodium, and see that it also burns with the formation of white fumes. Dissolve these fumes in water, and see that a solution of salt results.

viii. **Solubility of chlorine.**—Pass the gas through water for a short time, and observe that it is slightly soluble, the solution, known as chlorine water, smelling of the gas.

Preparation of chlorine.—It has been found that hydrochloric acid gas is a compound of hydrogen with some other constituent. To obtain this constituent free, we must evidently act upon the acid with something that will take away, or unite with, the hydrogen. Oxygen would immediately suggest itself if we could use it. Copper oxide contains oxygen, and we have seen that this reacts with hydrochloric acid, and that the oxygen unites with the hydrogen of the acid. The other constituent is not liberated, however, but unites with the copper to form a green powder

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[1] The gas is best collected in the open air or in a draught cupboard, as the fumes are very irritating if breathed.

(copper chloride). Some oxides, called *peroxides*, contain what we may term an extra amount of oxygen, and if these act on the hydrochloric acid, the quantity of the required constituent liberated by the union of the oxygen (of the oxide) with the hydrogen (of the hydrochloric acid) is more than sufficient to combine with the metal originally united to the oxygen. Some of this constituent, or chlorine, is hence liberated in the free state and we may prepare it in this way.

The plan adopted is to heat gently a solution of hydrochloric acid gas in water with black oxide of manganese, which is a peroxide, when chlorine is given off in large quantities in the form of a greenish-yellow gas.

The method used to get chlorine from common salt is first to mix it with black oxide of manganese, and then to heat the mixture with strong sulphuric acid, when chlorine is evolved as in the previous case. This process is really the same as the previous one, except that instead of first preparing hydrochloric acid from common salt and strong sulphuric acid, and then acting upon it with black oxide of manganese, the two experiments are combined. The three materials are heated together, and the hydrochloric acid as it is formed is decomposed by the manganese dioxide which is present.

Chlorine is heavier than the air and is usually collected by downward displacement, though it is sometimes collected over a strong solution of salt, or over warm water.

Properties of chlorine.—Chlorine is a gas with a greenish-yellow colour, from which fact [1] it gets its name. It has a disagreeable smell, and the gas, if breathed, causes distressing symptoms, which have been described as like those of an exaggerated cold in the head. The gas is soluble in water, and being heavier than air is usually collected by downward displacement. Its chief characteristic is its power of bleaching moist, vegetable colours. What happens

[1] The Greek word χλωρὸς (chloros), greenish-yellow.

is that the chlorine combines with the hydrogen of the moisture (which must be present for successful bleaching) to form hydrochloric acid and liberates the oxygen. This oxygen unites with the colouring matter to form a new chemical compound *which has no colour*, or, as chemists say, the oxygen oxidises the colouring matter.

The ease with which chlorine combines with hydrogen is seen not only by its action upon moisture, but in other ways. A lighted candle will continue to burn when plunged into chlorine, though with a very smoky flame. A candle is composed of hydrogen and carbon, and the flame continues, though with diminished brightness, because of the heat generated by the combination of the chlorine gas with the hydrogen of the candle to form hydrochloric acid gas. The carbon set free in the process is deposited as soot. The same explanation holds true for the spontaneous combustion of warm turpentine in chlorine gas.

Chlorine readily combines with metals to form chlorides. If finely divided iron, copper, antimony, and other metals be sprinkled into dry chlorine gas they at once combine with it, the heat of combination being sufficient to cause them to inflame. This happens more readily if the metals are first warmed. Sodium will burn in chlorine, and the sodium chloride which is formed may be readily recognised as common salt. Phosphorus, too, will continue to burn in chloride forming white clouds of a chloride of phosphorus.

66. QUANTITATIVE COMPOSITION OF HYDRO-CHLORIC ACID GAS.

i. **Density of hydrochloric acid and chlorine.**—Determine the density of well-dried hydrochloric acid gas and of chlorine in the manner described in Expt. 55. Take considerable care in doing the experiments, as accurate results are necessary, and can be obtained only with reasonable care.

Composition of hydrochloric acid.—-The density of the

hydrochloric acid gas, or, as we may call it, of *hydrogen chloride*, will be found to be about 0·00164 gram per c.c., and that of chlorine about 0·00319 gram per c.c. A litre of hydrogen chloride therefore weighs 1·64 grams, and two litres weigh 3·28 grams. We have already found, Expt. 64, vi. (*b*), that hydrogen chloride, or hydrochloric acid gas, contains one half its volume of hydrogen, so that two litres of the gas contain one litre of hydrogen, the weight of which is 0·09 gram. If, therefore, we subtract this from the weight of the two litres of hydrogen chloride we shall obtain the weight of chlorine present in these two litres. The result so obtained, 3·28 − 0·09 grams, is 3·19 grams, which is the weight of *one litre of chlorine.* It is proved, therefore, that hydrogen chloride consists of one half its volume of hydrogen and one half its volume of chlorine, or that 1 volume of hydrogen and 1 volume of chlorine are combined in 2 volumes of hydrogen chloride or hydrochloric acid gas

67. ACTION OF SULPHURIC ACID ON NITRE.

i. **Examination of nitre.**—(*a*) Again examine nitre according to the scheme on p. 13 ; or, carefully revise your notes of Expt. 14. Be careful to observe more fully the effect of heat on nitre ; when the nitre is strongly heated test the gas evolved with a glowing splinter, and into melted nitre drop a small piece of wood. Observe that it burns furiously. Nitre gives up oxygen very readily. It is a powerful *oxidising agent.*

(*b*) Place a little nitre in a test-tube, add some strong sulphuric acid, and warm. Notice that drops of a liquid condense at the top of the tube and run back.

ii. **Action of sulphuric acid on nitre.**—Into a stoppered retort, such as is used in the distillation of water (p. 52), place 30 or 40 grams of small crystals of nitre. Using a funnel, carefully introduce enough strong sulphuric acid to cover the nitre. Replace the stopper. Place the retort on a stand as shown in Fig. 79, and insert its neck in that of a flask, which is continu- ally kept cool by water, just as in the distillation of water. Gently heat the retort. Brown fumes are given off in abund- ance, and soon drops of a light, yellow liquid are seen to fall into the receiving flask. When enough liquid has distilled over, remove the laboratory burner, and while the materials in the

retort are still liquid, pour them, after removing the stopper, from the retort into an evaporating dish.
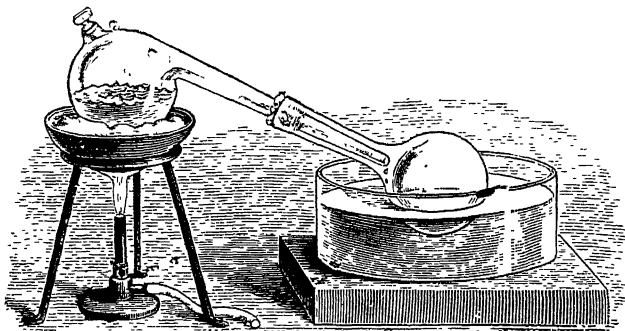
FIG. 79 —Heating sulphuric acid with nitre.

iii. **Properties of the liquid.**—(*a*) Pour a drop or two of the liquid from the receiving flask into a small glass of water. Taste a drop of the water by dipping the tip of your finger into it and then applying the finger to your tongue.

(*b*) Test the water with a blue litmus paper. The paper is turned red. The liquid is an acid.

(*c*) Pour a drop or two of the liquid from the receiving flask on to a fragment of copper in a test tube. Notice the brown fumes and the blue solution formed.

(*d*) Try the action of a little of the acid from the flask on a piece of cloth, and of a single drop on the palm of the hand. Notice the cloth is slowly destroyed and the skin is turned yellow.

iv. **The liquid produced is nitric acid.**—Compare carefully the liquid which you have obtained with the nitric acid in the laboratory bottles. Pay especial attention to the action of copper, which is very characteristic. Observe that your liquid is identical with nitric acid.

Nitric acid.—When sulphuric acid is heated with nitre, or saltpetre as it is also called, a liquid distills over, which, on examination, is found to be identical with the acid known as *nitric acid*, also sometimes called *aqua fortis*.

Nitric acid has certain properties associated with all acids. It has a sour taste and turns a blue litmus paper red. When pure, it is colourless and gives off colourless fumes.

But it is easily decomposed. When heated it gives off brown fumes. This happens slowly when the pure acid is placed in sunlight, a fact which accounts for the brown fumes above the liquid in bottles of nitric acid which have been kept for some time.

It is a very powerful acid, hence its name, *aqua fortis*, which dates from the time when all liquids were considered to be some kind of water. It destroys organic materials, like wood and cloth. It acts violently on most metals, like copper in Expt. 67, iii. (*b*). Gold and platinum are exceptions, for nitric acid has no action upon them. It very readily gives up some of the oxygen, of which it contains large quantities. Its great activity is due to the readiness with which it gives oxygen, and for this reason it is called an *oxidiser*.

QUESTIONS ON CHAP. XVII.

1. How is hydrochloric acid obtained? Give a short account of its chief properties.

2. You are provided with some nitre and strong sulphuric acid, describe fully how you would proceed to prepare a specimen of nitric acid.

3. Briefly indicate the reasoning which leads to the supposition that hydrochloric acid gas contains hydrogen united with another gas, and state how this second gas may be obtained from the acid.

4. What is the general effect of hydrochloric acid upon (*a*) metals, (*b*) oxides?

5. How may it be proved that hydrochloric acid gas consists of one half its volume of (*a*) hydrogen and (*b*) chlorine?

6. Describe the properties of chlorine, and state how you would obtain the gas from salt and then re-convert it into salt.

7. Under what conditions does chlorine unite with (*a*) hydrogen, (*b*) phosphorus, (*c*) sodium?

8. A lighted taper is placed in a jar of chlorine; what happens, and why?

9. How may chlorine be (*a*) obtained from, (*b*) converted into hydrochloric acid?

10. What is the action of sulphuric acid upon salt? What are the properties of both *products*?

11. From hydrochloric acid how could you obtain (*a*) hydrogen, (*b*) common salt?

12. By what tests would you ascertain whether effervescence caused by pouring sulphuric acid on a powder was due to the escape of carbon dioxide, hydrochloric acid gas, or hydrogen?

CHAPTER XVIII.

BASES, ACIDS, AND SALTS.

68. BASES.

i. **Caustic soda.**—(*a*) Place a small piece of sodium in an evaporating dish containing water. An alkaline solution of sodium hydrate, or caustic soda, which will consequently turn red litmus blue, is thus produced. Observe the peculiar soapy feel of the solution. Evaporate the solution to dryness, and examine the product : which is a white amorphous solid. Leave a little dry solid in an evaporating basin, and observe that it very soon becomes covered with a film of moisture—it is deliquescent. Compare it with the solid caustic soda in the laboratory and verify the statement that they are identical.

(*b*) Burn some sodium in a jar of oxygen or air (Expt. 38, x.), and examine the product formed. Observe that it is very soluble, yielding a solution with a soapy feel and an alkaline reaction. Evaporate the solution to dryness, the white solid formed is *caustic soda.*

ii. **Caustic potash.**—Repeat Expt. 68, i. *a*, using potassium instead of sodium. Observe that the reaction is far more vigorous. Evaporate the solution to dryness, and compare the product formed with the caustic soda, noting the great resemblance in properties. This substance is called *caustic potash.*

iii. **Action of aluminium on caustic soda.**—Place a strong solution of caustic soda in a test-tube, add some sheet aluminium and warm. Observe the evolution of hydrogen.

Properties of caustic soda and caustic potash.—These two compounds are white, deliquescent solids formed by the action of sodium and potassium respectively on water. They absorb moisture very readily and are both

powerful *alkalis*. In their general properties they resemble each other very closely, and, indeed, it might be thought that the two compounds are identical, but this is not so, as some of the later experiments will prove. When sodium is burnt in air an alkaline substance is also obtained, and it may therefore be thought that caustic soda is the oxide of sodium. This is not, however, quite correct. Caustic soda is formed by the action of sodium on water; if the sodium replaced all the hydrogen it is clear that the compound formed would consist solely of sodium and oxygen. But the sodium cannot replace the whole of the hydrogen, it only replaces one half, so that the resulting compound—the caustic soda—still contains some hydrogen. It consists of sodium, oxygen, and hydrogen, and when warmed with aluminium this hydrogen can also be displaced and its presence in caustic soda verified, as although the hydrogen evolved might have come from the water of the solution, yet, hydrogen is also evolved when aluminium acts upon dry, fused, caustic soda.

Hydroxides.—Compounds like caustic soda and caustic potash, which can be looked upon as obtained from water by replacing *a part* of the hydrogen by a metal, are termed *hydroxides*. Caustic soda may thus be called *sodium hydroxide*, and caustic potash in like manner *potassium hydroxide*.

69. FORMATION OF SALTS.

i. **Action of acids on metals, oxides, and hydroxides.**—(*a*) Act on some magnesium with sulphuric acid. Observe the evolution of hydrogen. Evaporate the solution *nearly* to dryness and allow it to crystallise. Examine the crystals formed.

(*b*) Burn some magnesium, and collect the white solid formed. Observe that it is a white, alkaline powder nearly insoluble in water, which glows when strongly heated, but does not undergo any chemical change. To some of this oxide add some diluted sulphuric acid, but not quite enough to dissolve all the oxide. Observe that the oxide dissolves *without* the evolution of gas.

Filter, partially evaporate, and allow it to crystallise. Compare the crystals with those formed in the previous experiment.

(*c*) Warm an excess of copper oxide with dilute sulphuric acid. Filter and, as before, obtain crystals of the product formed.

(*d*) Make a solution of sodium hydroxide by dissolving caustic soda in water. To a portion of the solution, in an evaporating basin, add dilute hydrochloric acid, drop by drop, until the solution has no effect upon either a red or blue litmus paper. The solution is then said to be *neutral*. Gently evaporate the solution, on a sand bath, until a dry, white residue is left. Then, by tasting the solid, satisfy yourself that it is common salt.

(*e*) Repeat the last experiment, substituting dilute sulphuric acid.

(*f*) Similarly perform the experiment, using dilute nitric acid.

(*g*) In a similar manner act on potassium hydroxide with hydrochloric acid, and observe that the product formed, a white solid, is *not salt*, although its taste recalls that of salt.

(*h*) Wash each of the products, obtained as above, with a little water, and then note that they have no action on litmus paper. They are *neutral compounds*.

Salts.—When a metal acts on an acid, the most general result is that the hydrogen contained in the acid is set free. In some cases, the hydrogen may react again with more of the acid and hence not be set free, but the usual result may be regarded as that just stated. The metal combines with the remainder of the acid and forms a new product, which we call a *salt*. We may, therefore, define a salt as the *product formed by replacing the hydrogen of an acid by a metal.*

Acids might be defined, therefore, as *substances which contain hydrogen capable of replacement by a metal,* but this definition is hardly sufficient; they are usually characterised by a sour taste and by their power of turning blue litmus red.

If, instead of a metal, the oxide of the metal acts on the acid, the salt is obtained as before, but hydrogen is not liberated, the reason being that the hydrogen, which has been replaced, combines with the oxygen of the metallic oxide to form water. Similarly with hydroxides, the hydrogen replaced unites with the oxygen and hydrogen

previously combined with the metal and forms water. Thus
  Metal and **acid** form a salt and **hydrogen.**
  Metallic oxide and **acid** form a salt and **water.**
  Metallic hydroxide and **acid** form a salt and **water.**

These metallic oxides, or hydroxides, are known as **bases,**
so that we could combine the last two statements and write
  A base and **acid** form a **salt** and **water.**

It must be remembered that *bases* are not necessarily
alkaline ; thus, copper oxide is a base, but has no action on
litmus paper. Only those bases which dissolve in water
give an alkaline reaction, and we may say that *an alkali* is
a *soluble base.*

How salts are named.—Salts are named from the
metal and acid which enter into their formation. Thus,
those containing copper are copper salts ; those containing
sodium are sodium salts, etc. Those formed from sulphuric
acid are termed *sulphates* ; those from hydrochloric acid
(hydrogen chloride), *chlorides* ; and those from nitric acid,
nitrates. The salts produced in Expt. 69 are therefore
named as follows : that produced from magnesium and
sulphuric acid is *magnesium sulphate* ; that formed from
copper oxide and sulphuric acid is *copper sulphate* ; that
formed from sodium hydroxide and hydrochloric acid is
sodium chloride ; that formed from sodium hydroxide and
nitric acid is *sodium nitrate,* etc.

70. AMMONIA AND AMMONIUM SALTS.

Certain other compounds, however, besides the metallic
oxides, may act as bases ; thus, we may obtain salts from
ammonia, which, by testing with red litmus, is seen to be
an alkali.

i. **Preparation of salts from ammonia.**—(*a*) Prepare salts from the
three acids, using the ammonia liquid of the laboratory in place
of caustic soda, in Expt. 69.

(*b*) Examine the solids formed, observing carefully the effect
of heat on each.

3. How would you obtain potassium nitrate and sodium chloride from their respective acids and bases?

4. What is ammonia? Give an account of its preparation and more characteristic properties.

5. In what respect does a solution of ammonia resemble a solution of caustic soda or caustic potash?

6. What are ammonium salts; how are they prepared?

7. Draw the apparatus, and describe how you would prepare ammonia gas from the ordinary solution of ammonia.

8. State the composition and principal properties of ammonia. What happens when it is mixed (a) with water, (b) with nitric acid?

9. How would you obtain crystallised sal-ammoniac from a solution of ammonia? How would you obtain the solution of ammonia from this product?

10. What are chlorides? How may they be obtained? Give examples.

11. If you were given some nitric acid and some caustic potash, describe *exactly* how you would proceed in order to obtain large crystals of nitre.

12. What are hydroxides? Give examples of metallic hydroxides. What products are formed when these hydroxides react with acids?

13. Write the names and describe the appearances of the substances produced when diluted sulphuric and hydrochloric acids respectively are mixed with soda and lime respectively, and the water evaporated away.

14. Describe those properties possessed by sulphuric, hydrochloric, and nitric acids, which lead to these substances being designated ' *acids.*'

15. Describe the terms acid, base, salt, and indicate how they are mutually related.

16. What are sulphates? Describe a method of preparation, and give a *short* account of the appearance and properties of three sulphates.

17. What is the difference between the action of an acid, say hydrochloric acid, upon iron and oxide of iron?

18. What is meant by an alkali? Compare the chemical behaviour of an alkali with that of an acid. Describe the appearance of any two alkalis.

19. Describe exactly how you would prepare a cylinder of dry ammonia gas, and also a strong solution of hydrochloric acid.

ADDITIONAL PRACTICAL EXERCISES.

1. Examine and describe as in Chap. ii. the following: soil, lime-stone, aluminium, sugar, wood, clay, alum, borax, black lead.

2. Find out whether an egg undergoes any change of weight on being boiled.

3. Find out whether wax changes in weight when melted. Does it change in density, and if so, is it more or less dense when liquid ?

4. Determine the quantity of sand (or other insoluble solid) in a mixture of sand and salt.

[NOTE.—Weigh out a quantity, from 1 to 2 grams, of the mixture and shake well with sufficient water. When all the soluble solid is considered to be dissolved filter and wash the filter paper with the insoluble residue two or three times, by pouring distilled water through it. Then place the funnel and filter paper in a steam oven (or other oven kept at about 100° C.) to dry, placing also another similar filter paper moistened with distilled water. After remaining for a couple of hours weigh the filter paper and residue, using the second filter paper as a counterpoise on the other scale-pan. Calculate percentage of sand.]

5. Determine the weight of soluble and of insoluble solid in the mixtures provided and compare the sum of the weights obtained with the total used.

6. Determine the quantity of solid matter dissolved per litre of ordinary tap water.

7. Determine the quantity of solid matter dissolved in sea water per litre.

8. Determine as in (4) the quantity of soluble and insoluble matter in soil.

9. Examine samples of moist sugar for insoluble matter, if present, and determine its percentage.

10. Determine the percentage of ash in coal.

[NOTE.—Weigh a crucible ; then again weigh the crucible after placing in it about 2 grams of coal. Heat strongly until only a white ash remains, and again weigh when cool.]

11. Determine the percentage of ash in wood, charcoal, and paper.

12. Determine the percentage of sal-ammoniac in a mixture of sal-ammoniac and salt.

[Think carefully of the properties of the two substances, and hence find a method for their separation.]

13. Determine the percentage of sulphur in a mixture of sulphur and powdered glass :

(1) By making use of the action of heat on sulphur.

(2) By making use of its solubility in certain liquids.   Compare the two results.

14. Determine by 2 methods the weight of sal-ammoniac and powdered glass in a mixture, and compare the results.

15. Make up a solution containing exactly 100 gms. of salt per litre, and determine the concentration by evaporation of 25 c.c.

16. Weigh out sufficient magnesium to displace exactly 200 c.c. of hydrogen.   Determine experimentally the volume actually displaced.

17. Determine the quantity of zinc in an amalgam of zinc and mercury.

[The hydrogen evolved by the action of dilute sulphuric acid may be considered to be due only to the zinc.]

18. Determine the quantity of oxygen and nitrogen in the ' air ' dissolved in water.

19. Determine the percentage of lead and oxygen in litharge by heating the litharge in a current of hydrogen (coal gas may be employed) in a hard glass tube.

20. Determine the solubility of lime.

21. Determine the percentage of potassium chlorate in a given mixture of potassium chlorate and manganese dioxide (1) by making use of the solubility of the former, (2) by a determination of the weight or volume of the oxygen evolved on heating.

22. Determine the quantitative composition of oxide of mercury by finding the loss of weight on heating.

[Take care to adopt adequate precautions to avoid loss of mercury by volatilisation.]

23. Determine the percentage of potassium chloride in a mixture of potassium chlorate and chloride.

24. Determine the percentage of chalk in a mixture of chalk and lime by (1) heating, (2) action of an acid.

25. Determine the solubility of lime in distilled water at ordinary temperatures.

26. Determine the quantity of water of crystallisation in soda crystals.

27. Determine the percentage of carbon dioxide in a sample of limestone ; and also the quantity of material insoluble in acid.   (The latter by the action of excess of acid, filtration, and weighing the dried filter paper and residue as in Ex. 4.)

28. Determine the weight and volume of hydrogen evolved by the action of 1 gram of aluminium on acids, and calculate the weight of aluminium necessary to displace 1 gram of hydrogen.

29. Determine the weights of copper and of oxygen in copper oxide.

[By heating in a current of hydrogen until completely converted into copper. (**N.B.—Precautions always necessary in work with hydrogen,** see p. 124.)]

30. Weigh out in a crucible sufficient chalk to yield exactly 1 gram of pure lime. Heat very strongly, and find experimentally the weight of lime produced.

31. Determine the percentage of carbon dioxide in a given mixture of (1) carbon dioxide and oxygen, (2) carbon dioxide and nitrogen.

[Think carefully of the properties of carbon dioxide, and find a method by which it can be separated from either of the other gases.]

32. Determine the percentage of hydrogen in ordinary concentrated sulphuric acid.

[Proceed as in Exp. 48, but use a weighed quantity of the sulphuric acid and employ *excess* of zinc.]

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